50

A SHORT COURSE

IN

QUALITATIVE ANALYSIS

BY

J. M. CRAFTS,

FORMERLY PROFESSOR OF GENERAL CHEMISTRY IN THE CORNELL UNIVERSITY.

THIRD EDITION.

REVISED BY

CHARLES A. SCHAEFFER, A.M., PH.D.,

PROFESSOR OF GENERAL AND ANALYTICAL CHEMISTRY IN
THE CORNELL UNIVERSITY.

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This Book is inscribed by the Author

TO

M. ADOLPH WURTZ,

AS A TOKEN OF AFFECTIONATE REGARD, TO A FRIEND AND TEACHER,
AND AS A TRIBUTE OF RESPECT AND ADMIRATION, TO THE
MASTER OF A SCHOOL IN MODERN CHEMISTRY.

PREFACE TO THE THIRD EDITION.

IN preparing a new edition of this work, the editor has made no essential alteration of the plan adopted by Professor Crafts in the previous editions. He has endeavored, however, as far as possible, to embody the practical results of the great advance made in both theoretical and practical chemistry within the past ten years. The few important changes thereby rendered necessary may be briefly specified.

Chapter II., on chemical nomenclature, has been entirely rewritten. The usefulness of the work has been somewhat extended by the introduction of several elements not included in the previous editions, viz. : strontium, cadmium, iodine, and bromine.

Further changes have been made in the methods of separation and detection of several of the elements ; more reliable processes being substituted for those which in the hands of students were found not to be entirely satisfactory.

The editor desires especially to acknowledge his obligation to his friend and colleague, Dr. G. C. Caldwell, for numerous valuable suggestions of which he has availed himself.

ITHACA, December 17, 1879.

PREFACE TO THE FIRST EDITION.

This little work was written for the use of a class of students in the Cornell University, who take a year's course of chemistry, including four hours a week of laboratory practice; reference was also had to the requirements of the scientific students in Union College, whose course is nearly equivalent to that mentioned. The author is indebted to the kindness of Professor Perkins for many valuable suggestions, and for the compilation of Tables IV. and V. at the end of the book.

A considerable portion of the introductory part of this book is devoted to an explanation of the theory of chemical reactions and nomenclature. Many of the standard works on analytical chemistry still use the old notation, and the formulas to be found in them do not correspond to those used in the best text-books on general chemistry; and none of the elementary works for the laboratory-student supply the want, often felt by him, of a system of rules, at hand for use at the moment when he most requires them, namely, when he is writing the formula of a reaction at his desk in the laboratory with his tests before him. The author has had these points in view in writing the two introductory chapters.

It might be objected that the theory of chemical notation should be found in text-books on general chemistry; but even when the student has mastered the rudiments of the science, as given in any of the best modern works, he will find the arrangement of many of them inconvenient for reference, although it is excellent for instruction, and, moreover, it is by no means necessary that the study of those works should precede that of analytical chemistry.

It is quite feasible to commence a course of experiments in the laboratory concurrently with the study of any of the elementary text-books on general chemistry, or with the attendance upon lectures, illustrated with the usual experiments. Chemistry certainly becomes a more attractive study when the practical and the theoretical present themselves side by side, so that while the theory explains the experiment, the experiment awakens an interest in the theory ; and no course of study is more apt to interest the beginner in chemistry than that of the admirably simple and delicate tests of qualitative analysis ; tests which illustrate the general laws of the science, while they have a very direct bearing upon some of the problems of every-day life.

Analytical chemistry, besides its immediate value as an important branch of knowledge, cannot be too highly prized as affording a convenient introduction to the methods of investigation used in an experimental science, and as offering a means of education of many faculties, which are not easily developed by school or university training. The importance of laboratory experiments is awakening every day increased attention, and the time is fast passing by when chemistry is taught to persons, who suppose that they have a vocation for a scientific profession, only by lectures and recitations.

The system of analysis, given in Part III., is founded upon that of Fresenius, and includes a minute description of all the steps to be taken in performing tests and directions for passing from one test to another; and since these details, which are required by the beginner, not only become unnecessary, after a certain familiarity with analytical work has been acquired, but also render the scheme less convenient for reference, tables similar to those of Will have been added, which indicate the important tests and leave the experience of the student to suggest the proper mode of applying them.

The Tables IV. and V. are intended to record a number of facts in analytical chemistry in a compact form, and to give an exact conception of what is meant by the insolubility of a precipitate, and a means of judging of the advantages of different methods of precipitation.

ITHACA, August 19, 1869.

PREFACE TO THE SECOND EDITION.

A FEW changes have been made in the second edition, which have been suggested by the use of the previous one in the laboratory and some practical directions have been added, which are intended to mark out more closely the course of study which it seems most advisable to follow.

One of the principal applications of analytical chemistry is to the investigation of minerals, and it may not be out of place here to recommend to the consideration of instructors and students the combination of mineralogical work with laboratory practice. A convenient order of study is, first to acquire familiarity with the methods of detection and separation of bodies in solution, which are given in Part III., and then, before taking up the analysis of solid bodies, to study their crystallographical and general mineralogical characters from any elementary work treating of the subject. Turning again to the general analysis of solid bodies (minerals and technical products of all kinds), the course should be made as varied and extended as possible ; and it will be found that the drill in blowpipe practice and the study of the physical characters of minerals make an excellent preparation for the preliminary testing of solids, Part III., and often afford hints which are of great value in the subsequent performance of the complete analysis in the wet way.

It is quite possible for a class to complete a course like the one proposed by working four hours a week for a year in the laboratory.

ITHACA, 1870.

A SHORT COURSE
IN
ANALYTICAL CHEMISTRY.

PART I.—INTRODUCTION.

CHAPTER I

THE analytical chemist is not generally required to investigate a large range of substances, and the theory of the composition of those bodies which he usually deals with can be briefly explained.

Elements. The ultimate result of the analysis or decomposition of all matter is the discovery of a number of substances which cannot be decomposed, and are therefore called simple or elementary substances. The table on the next page contains the names of thirty-six elements. They are chosen from among the seventy or more elements which have been discovered, because they are more frequently met with than the others, and because they are the only ones treated of in this work.*

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* Within the past four years ten new elements have been reported by their discoverers. How many of these will be substantiated it is at present impossible to say. On the other hand, very recent investigations seem to show that at least several of our well-known elements are really compound bodies. The exact number of elements is, therefore, especially at the present crisis, very difficult to fix.

	MONADS.	DYADS.	TRIADS.	TETRADS.	PENTADS.	HEXADS.
NON-METALLIC ELEMENTS.	Fluorine, F.. 19.	Oxygen, O...... 16.	Nitrogen, N.. 14.		Nitrogen, N.. 14.	Sulphur, S.... 32.
	Chlorine, Cl.. 35.5	Sulphur, S...... 32.	Boron, B..... 11.		Phosphorus, P. 31.	
	Bromine, Br.. 80.					
	Iodine, I.....127.			Silicon, Si...... 28.		
	Hydrogen, H. 1.			Carbon, C...... 12.		
METALS.			Arsenic, As... 75.	Platinum, Pt....197.4	Arsenic, As... 75.	
			Antimony, Sb.122.		Antimony, Sb..122.	
			Bismuth, Bi...208.		Bismuth, Bi...208.	
		Lead, Pb.......207.		Tin, Sn........118.		
	Silver, Ag...108.	Copper, Cu..... 63.4	Gold, Au.....196.			
		Mercury, Hg....200.				
		Cadmium, Cd...112.				
		Nickel, Ni...... 59.				
		Cobalt, Co...... 59.				
		Iron, Fe........ 56.		Iron, Fe........ 56.		
		Manganese, Mn.. 55.		Manganese, Mn. 55.		Manganese, Mn. 55.
		Zinc, Zn........ 65.		Chromium, Cr... 52.		Chromium, Cr.. 52.
				Aluminium, Al.. 27.4		
		Magnesium, Mg. 24.				
		Calcium, Ca..... 40.				
		Strontium, Sr.... 87.6				
	Sodium, Na.. 23.	Barium, Ba.....137.				
	Potassium, K. 39.					

The task of the analytical chemist is to recognize these elements when they occur alone, or to separate them from the compounds of which they form a part, or most frequently to resolve a compound substance into simpler compounds, and to isolate the latter in such form that they can be easily recognized. When the different elements are known which constitute such simpler compounds, the composition of the complex body from which they were obtained can be determined.

All compound substances are formed by the union of chemical elements, and a previous study of the manner in which the elements combine with each other is essential to the successful pursuance of analytical investigations.

The language in which chemists express their ideas regarding the chemical constitution of bodies comprises certain symbols or abbreviations, to which a conventional meaning is attached, and formulas, which are made by grouping symbols together.

Symbols. The symbols which stand in the table after the names of the elements are abbreviations, which are used instead of the names in writing chemical formulas.

Chemical Formulas describe by means of symbols the chemical constitution of bodies.

Example:—HCl is the formula of hydrochloric acid, and signifies that it is composed of hydrogen and chlorine.

The Combining Weights are the numbers standing in the table after the symbols. They are peculiar to each element, and denote the proportion by weight in which it unites with other elements. In formulas the symbols stand for these weights, as well as for the names of the elements; thus, in the formula of hydrochloric acid, HCl signifies that 1 part by weight of hydrogen is combined with 35.5 parts by weight of chlorine.

ELEMENTS COMBINE WITH EACH OTHER IN NO OTHER PROPORTIONS BY WEIGHT THAN THOSE EXPRESSED BY THE

COMBINING NUMBERS (OR BY VERY SIMPLE MULTIPLES OF THEM).

This statement comprises the law of definite proportions and also that of multiple proportions.

USUALLY THE MULTIPLES OF THE COMBINING NUMBERS, 2, 3, 4, 5, 6, 7, EXPRESS THE PROPORTIONS IN WHICH THE ELEMENTS COMBINE WITH EACH OTHER.

Example:—Nitrogen combines with oxygen only in the proportions expressed by the formulas: N_2O, NO, N_2O_3, NO_2, and N_2O_5, or, referring to the table for the combining numbers for which the symbols stand, the following proportions appear:

2 × 14	parts	N	unite	with	16	parts	O.
14	"	"	"	"	16	"	"
2 × 14	"	"	"	"	3 × 16	"	"
14	"	"	"	"	2 × 16	"	"
2 × 14	"	"	"	"	5 × 16	"	"

Chemical Atoms. It is supposed that the utmost limit to which the division of matter could be carried would lead to its separation into a great number of particles, so small as to be incapable of further division. With reference to their quality of indivisibility, such particles of matter are called atoms (from ἀ, privative, and τέμνω, I cut). Atoms, therefore, are the indivisible constituents of matter. It is further supposed that chemical combination consists in the union of atoms, or groups of atoms, and chemical decomposition in the separation of atoms, or groups of atoms; and a chemical change supposes a change in the arrangement or grouping of the atoms of a body involving the destruction of the previous arrangement.

Atomic Weights. The atomic theory attaches a new meaning to the combining weights of the elements, and defines them as the relative weights of atoms. Thus, if the weight of an atom of hydrogen is 1 × w (w = a very small unknown quantity), the weight of an atom of chlorine is 35.5 × w. Not the absolute weights of atoms, but their relative weights, have been discovered. The weight of an atom of hydrogen is taken as a standard, and called 1; hence the weight of the atoms of other elements are expressed in terms of this unit. Example, 35.5 for chlorine, 16 for oxygen. With reference to the above theory, the combining weights are usually called atomic weights. The full meaning, therefore, of the chemical formula HCl is, that the body which it represents consists of compound atoms, each one containing an atom of hydrogen and an atom of chlorine. An atom of chlorine weighs 35.5 times as much as an atom of hydrogen, so that the proportion by weight of each constituent of the body is expressed by its formula.

Chemical Notation. Any change in the constitution of bodies, as well as their formation and decomposition, involves what is called a chemical reaction.

A CHEMICAL REACTION may be described as a change in the arrangement, or the state of combination of the atoms of bodies. Such a change can be denoted by combining formulas together in the same way that quantities are combined in common algebraic calculations. The signs used are +, —, and =. Coefficients are only used to multiply the symbols to which they are joined. When placed on the line, they multiply all the symbols which follow. When placed below or above the line, they are used to multiply only the symbol, or the group of symbols in brackets, that immediately precedes them. Brackets are used to distinguish certain groups of atoms in a compound from the remaining atoms. The combination of atoms is expressed by writing their symbols side by side, or by grouping them together without + or —. The sign + expresses that the bodies connected by it are brought in contact with each other by addition, but that they are not combined.

Example :—$Ba(NO_3)_2 + CaSO_4$. The formula indicates that a compound, $Ba(NO_3)_2$, containing barium, nitrogen, and oxygen, in the proportion of 1 atom of barium combined with twice 1 atom of nitrogen and twice 3 atoms of oxygen, is brought in contact with a compound [$CaSO_4$] containing calcium, sulphur, and oxygen, in the proportion of 1 atom of calcium combined with 1 atom of sulphur and 4 atoms of oxygen. A reaction is denoted by combining the formulas of the bodies which take part in it, as in an ordinary equation. The formulas before the sign (=) indicate the state of combination of the atoms before the reaction; those after the sign of equality (=) show the state of combination of the atoms after the reaction.

Example:—$Ba(NO_3)_2 + CaSO_4 = BaSO_4 + Ca(NO_3)_2$.

The equation expresses the result of bringing in contact the bodies described above, viz.: the formation of new compounds containing barium, sulphur, and oxygen, and calcium, nitrogen, and oxygen.

CHEMICAL AFFINITY.

The force which impels atoms to unite with other atoms is called chemical affinity. The quantity or the nature of the force inherent in the atoms of every substance determines the chemical properties of the substance. The study of the results of the action of chemical affinity is the province of chemistry.

The phenomena which the action of chemical affinity gives rise to can best be studied under several heads.

Firstly.—Chemical affinity may cause the atoms of an elementary body to unite among themselves. Only the cases of such action in which the element is capable of assuming the simplest physical condition of matter, namely, the form of a gas, have been studied satisfactorily.

The following conclusions in regard to the state of combination of the atoms of elementary bodies have been arrived at:

The atoms of mercury and zinc remain separate in the gaseous state.

The atoms of hydrogen, oxygen, chlorine, bromine, iodine, nitrogen (and sulphur at a temperature higher than 1000° centigrade) are united in groups of two atoms.

The atoms of phosphorus and arsenic in the gaseous state unite in groups of four atoms. Sulphur, at a temperature of 500° centigrade, in groups of six atoms.

The formulas for these bodies in a gaseous state are:

Hg
Zn

HH or H_2
OO or O_2
ClCl or Cl_2
BrBr or Br_2
NN or N_2
SS or S_2

PPPP or P_4
AsAsAsAs or As_4

SSSSSS or S_6

Molecule. At this point a definition of the term molecule is required. A molecule is the smallest particle of a body which can exist alone.

The molecules of elementary bodies contain one or more atoms of the same kind. The molecules of compound bodies contain two or more atoms of different kinds.

A molecule of mercury, hydrogen, hydrochloric acid, or water is represented by the respective formulas, Hg, H_2, HCl, or H_2O.

Atom. An atom may be further defined as the smallest particle of matter which can take part in a chemical reaction. Atoms, therefore, appear while a chemical reaction is going on, although it is impossible to suppose that a physical subdivision of matter could be carried further than the isolation of molecules.

Thus in the reaction: $Na_2S + CuCl_2 = CuS + 2\,NaCl$, the force of chemical affinity breaks up the molecules (Na_2S) and

($CuCl_2$) to form the new ones (CuS) and ($NaCl$), and during this reaction the atoms Na, S, Cu, and Cl must be set free from their combinations, and therefore must exist as atoms.

Secondly.—Chemical affinity may combine atoms of a single element, or groups containing atoms of one or more elements, with the atoms of another element, or with groups containing atoms of one or more other elements.

Examples:—$Hg + Cl_2 = HgCl_2$.
$NH_3 + HCl = NH_4Cl$.

The inverse action frequently takes place through the agency of heat or of some other force, and groups of atoms (molecules) break up into other groups, which are usually simpler in constitution than the primitive ones.

Example:—$Hg(CN)_2$ when heated becomes $Hg + (CN)_2$.

Thirdly.—Chemical affinity may cause compound bodies, brought in contact with each other, to mutually exchange some of their constituents; or an atom or a group of atoms may substitute itself for another atom or for a group of atoms in a compound body.

Examples:—$Na_2CO_3 + BaCl_2 = BaCO_3 + 2NaCl$.
$CuSO_4 + Zn = ZnSO_4 + Cu$.

THE QUALITY OF THE CHEMICAL AFFINITY inherent in the atoms of each element determines the part which the element will play in the different chemical changes mentioned above. It is usually necessary to study each particular case, in order to determine the exact result of bringing in contact any two substances. Empyrical rules, however, defining the nature of the chemical affinity of the elements and the consequences of its action, can be given in a few cases. These rules are not capable of a very strict application, but they serve to indicate, in most cases, when a number of bodies are brought together in a reaction, those which will probably combine with each other. Gold is attacked by acids less readily than the metals of the arsenic group. The difference between the affinities of the metals which are ranged in the same group in the table (page 2) is too slight to be of consequence in the application of rule second.

First.—COMBINATION USUALLY OCCURS BETWEEN METALS AND NON-METALLIC ELEMENTS, LESS READILY BETWEEN DIFFERENT NON-METALLIC ELEMENTS, AND LEAST READILY BETWEEN DIFFERENT METALS. See table of the elements, page 2.

The following non-metallic elements, or oxygen compounds of non-metallic elements, combined with hydrogen form acids, and combined with metals form salts. STRONG ACIDS—sulphuric, H_2SO_4; nitric, HNO_3; chloric, $HClO_3$; chlorhydric, HCl. WEAK ACIDS—sulphurous, H_2SO_3; chromic, H_2CrO_4; phosphoric, H_3PO_4; boracic, H_3BO_3; oxalic, $H_2C_2O_4$; acetic, $HC_2H_3O_2$; fluorhydric, HF; sulphydric, H_2S; cyanhydric, HCN; carbonic, H_2CO_3; silicic, H_4SiO_4. A metal has a tendency to substitute itself for the hydrogen in a strong acid, to form a salt with it, in preference to a weak one; so that a strong acid usually displaces a weak one from its salts.

Second.—THE METALS WHICH STAND LOWEST IN THE TABLE (page 2) HAVE THE GREATEST TENDENCY TO COMBINE WITH THE STRONGEST ACIDS.

Example:—$CuSO_4 + Zn = ZnSO_4 + Cu$.

Third.—Another rule which sometimes takes precedence of the second is, that WHEN FROM SOME OF THE CONSTITUENTS OF DIFFERENT COMPOUNDS IN SOLUTION, AN INSOLUBLE BODY CAN BE FORMED, THE ELEMENTS WHICH WOULD COMPOSE SUCH A BODY GENERALLY UNITE WITH EACH OTHER.

Example:—$Na_2CO_3 + Ca(HO)_2 = 2NaHO + CaCO_3$. The compound $CaCO_3$ is formed in virtue of its insolubility when the reaction takes place in an aqueous solution.

QUANTIVALENCE.

The quality of the chemical affinity of the elements determines the nature of reactions. The quantity of their chemical affinity determines the proportions in which the elements combine with each other. The result of the combination of two or more atoms with each other is the neutralization of the chemical force which brings them together, so that usually no force is left in the compound tending to combine other bodies with it. Thus a molecule of HCl has no power to combine further with atoms of H or of Cl. The atoms of some elements are animated with greater quantities of chemical force than those of other elements. Thus an atom of oxygen may unite with an atom of hydrogen, and still be capable of combining with another atom of hydrogen, or with an atom of chlorine. If an atom of hydrogen has one unit of chemical force, an atom of oxygen has two units, carbon has four, and nitrogen has five. The number of units of chemical force residing in an atom is called its QUANTIVALENCE. The table of the elements, page 2, classifies them, according to their quantivalence, into monatomic elements, or monads; and diatomic elements, or dyads, etc.

Some of the units of chemical force of an element may lie dormant until developed by the approach of some other force, which awakens them. Iron, for instance, acts as a dyad when no element is present to call forth all of its four units of chemical force. It is worthy of note that in such cases two units of force generally disappear together, as if they became dormant by each one neutralizing the effect of the other. It will be noticed that in the table, page 2, sulphur is placed in the column of dyads, and also in that of the hexads; this is because in many cases four of the six units of chemical force in sulphur lie dormant, and the element plays the part of a dyad. In nickel, cobalt, and iron, two of the four units of chemical force frequently lie dormant, and hence these metals often act as dyads. Only the valence of each element, which is usually displayed in the kind of reactions which we have to consider, is shown in the table.

The knowledge of these facts is essential to enable the student to write formulas correctly.

Examples:—The compound of barium and chlorine must contain two atoms of chlorine, combined with one atom of barium. Its formula is $BaCl_2$. The greatest amount of oxygen that an atom of carbon can unite with is two atoms. The formula of the compound is CO_2. When zinc is substituted for silver in a compound, one atom of zinc, with its two units of chemical force or affinity, takes the place of two atoms of silver, because each atom of silver has only one unit of chemical force. Thus, in writing formulas, an atom of one element is equivalent to or takes the place of another element of the same class. In comparing elements of different classes, their value in an equation depends upon the number of units of chemical force which they contain.

TWO MONADS ARE EQUIVALENT TO A DYAD; THREE MONADS TO A TRIAD, ETC. THREE DYADS ARE EQUIVALENT TO TWO TRIADS, ETC.

Examples:—$PbO + 2HCl = PbCl_2 + H_2O$.

$SiCl_4 + 2H_2O = SiO_2 + 4HCl$.

QUANTIVALENCE OF GROUPS OF ATOMS.—When part of the affinities or units of chemical force of a polyatomic group are satisfied or neutralized, the residual valences, or those which remain free, determine the quantivalence of the group.

Example:—Nitrogen, a pentad, when combined with three atoms of hydrogen, is capable of uniting with two other monads or with one dyad—NH_3 can unite with H and Cl to make NH_4Cl. Moreover, certain groups of elements take part in many chemical reactions without being broken up, and, so far as this is the case, they may be considered as playing the part of elementary bodies in the reactions. Such groups are frequently called radicals, with reference to a theory that they are the roots of compounds.

The following table, showing the quantivalence of such groups, will be found convenient in writing formulas :

Monads.	Dyads.	Triads.	Tetrads.	Hexads.
HO*	SO_4‡	PO_4	SiO_4	Fe_2§
NO_3†	SO_3	BO_3	$Fe(CN)_6$	Cr_2
ClO_3	CrO_4			Al_2
$C_2H_3O_2$	CO_3			$Fe_2(CN)_{12}$
NH_4	C_2O_4			
CN	$C_4H_4O_6$			
	Hg_2			

Examples:—$Na(HO)$; $H(NO_3)$; $(NH_4)\,(HO)$; $Pb\,(C_2H_3O_2)_2$;

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* The group HO is a monad, because one of the two affinities of the oxygen atom is satisfied and nullified by the affinity of the hydrogen with which it is combined, leaving one affinity free in the oxygen atom.

† The existence of the group NO_3 might seem inconsistent with the five-atomic character of nitrogen. The explanation is, that four affinities in the nitrogen atom are used to unite to it two atoms of oxygen, and the fifth combines with only one of the affinities of a third atom of oxygen, leaving one oxygen affinity free, which determines the monatomic character of the group.

‡ A similar argument applies to the case of the group SO_4. Here, sulphur being hexatomic, four affinities are used to combine with those of two atoms of oxygen, while the two remaining affinities of the sulphur atom are each combined with one in each of the two remaining atoms of oxygen, leaving two oxygen-affinities free.

§ In the case of Fe_2, etc., two of the eight affinities, belonging to two atoms of iron, are used to bind the two atoms together, leaving six free, and for this reason the group Fe_2 is six-, and not eight-atomic. It may be noticed that groups of this nature, having free chemical affinities, can only be found in the case of polyatomic elements, because, when two monatomic elements combine with each other, no chemical affinity can be left free. For fuller explanations of the laws of chemical combination, see Frankland's Lecture Notes for Chemical Students.

$H_2(SO_4)$; $Ba\,(CrO_4)$; $Hg_2\,(Cl_2)$; $H_3\,(PO_4)$; Fe_2Cl_6 ; Cr_2O_3 ; $Al_2\,(HO)_6$.

IN WRITING FORMULAS, THE SYMBOL OF THE ELEMENT, OR GROUP OF ELEMENTS, HAVING MOST DECIDEDLY THE CHARACTER OF A METAL, IS PLACED FIRST.

Example :—$NaNH_4HPO_4$.

CHAPTER II.

CHEMICAL NOMENCLATURE.

When metallic sodium comes in contact with water a reaction takes place, as indicated by the following equation :

$$Na_2 + 2H_2O = 2NaHO + H_2.$$

Each molecule of water is broken up, setting free an atom of hydrogen and receiving in its place an atom of sodium. The new body thus formed may then be considered as a molecule of water in which one atom of hydrogen has been replaced by sodium. Or it may be considered as an atom of a monad metal united to the monovalent radical HO, known as hydroxyl. Furthermore, it may be taken as a type of a numerous class of bodies, known as Metallic Hydrates, all of which consist of one or more atoms of metal united to an equivalent amount of hydroxyl, as may be seen from the following

Examples:—KHO ; $Ca(HO)_2$; $Fe(HO)_2$; $Fe_2\ (HO)_6$; $Al_2\ (HO)_6$.

All of the metallic hydrates, when acted upon by a certain other class of bodies, are similarly affected, as can be shown by the following equations :

$$\begin{array}{llll}
HCl & + KHO & = KCl & + H_2O ; \\
2HCl & + Ca(HO)_2 & = CaCl_2 & + 2H_2O ; \\
6HCl & + Fe_2(HO)_6 & = Fe_2Cl_6 & + 6H_2O ; \\
HNO_3 & + KHO & = KNO_3 & + H_2O ; \\
6HNO_3 & + Fe_2(HO)_6 & = Fe_2(NO_3)_6 & + 6HO ; \\
H_2SO_4 & + KHO & = KHSO_4 & + H_2O ; \\
H_2SO_4 & + 2KHO & = K_2SO_4 & + 2H_2O ; \\
H_2SO_4 & + Ca(HO)_2 & = CaSO_4 & + 2H_2O.
\end{array}$$

By inspection of the above equations it will be seen that we have in the left-hand member of each a metallic hydrate, and in the right-hand member of each water. Furthermore the substance indicated first in each equation contains hydrogen, and when this hydrogen is replaced by the metal of the metallic hydrate the formula of the substance indicated first in the right-hand member of the equation is obtained.

The original substances containing hydrogen are called acids, and the compounds occurring first in the right-hand members of the above equations are salts.

The above we may formulate, then, in the following definitions:

Acid. An acid is a compound which, when brought into contact with a metallic hydrate, exchanges the whole or a part of its hydrogen for the metal of the metallic hydrate; water being formed at the same time.

The acids may be divided into two groups.

Firstly, Those which consist of hydrogen united directly to another non-metallic element as HCl; HBr; HF; H_2S.

Secondly, The oxygen acids, or those in which the hydrogen is united with another non-metal, or with a group, by means of oxygen, as in $HClO$; HNO_3; H_2SO_4. Here the arrangement of the chemical forces may be shown by the following graphic formulas: $H - O - Cl$; $H - O - N\ O_2$; $H_2 = O_2 = SO_2$.

Salt. A salt is a compound formed by replacing the hydrogen of an acid by a metal.

All compound bodies are susceptible of decomposition by a current of electricity, and when thus decomposed resolve themselves into a positive atom or group, and a negative atom or group. In writing the ordinary formulas of compounds the electro-positive atom or group is always placed first, as

$$\overset{+}{\mathrm{H}}\overset{-}{\mathrm{Cl}};\quad \overset{+}{\mathrm{H_2}}\overset{-}{\mathrm{O}};\quad \overset{+}{\mathrm{Na}}\overset{-}{(\mathrm{HO})};\quad \overset{+}{\mathrm{Ba}}\overset{-}{(\mathrm{SO_4})};\quad \overset{+}{(\mathrm{NH_4})_2}\overset{-}{(\mathrm{SO_4})}.$$

The names of chemical compounds, which are used in works on analytical chemistry, do not describe their constitution as fully as formulas do; they serve, however, to identify the bodies and to recall certain principles of classification.

In the ordinary operations of analysis, the only bodies which are dealt with are binary compounds (*i.e.*, compounds which only contain two elements, as KCl and H_2O) and ternary compounds, which contain a non-metallic element and oxygen combined with a metal or with hydrogen, as $BaSO_4$, H_3PO_4.

In the statement of the rules for the naming of compounds the subject will be divided under these two heads:

I. BINARY COMPOUNDS.

a. ACIDS.—Two distinct methods are in common use for naming binary acids. According to the one the name consists of *hydro* preceding the root of the name of the other element and terminates in *ic;* as, HCl, hydrochloric acid; HBr, hydrobromic acid; H_2S, hydrosulphuric acid. By the other method the characteristic element is indicated first, and the names of the three acids just mentioned would be, chlorhydric acid, bromhydric acid, and sulphydric acid.

b. ALL OTHER COMPOUNDS.—In determining the names of binary compounds not acids, if there exists but one compound of the two elements in question, the positive element takes the termination *ic* and the negative *ide;* as, NaCl, sodic chloride; K_2S, potassic sulphide; BaO, baric oxide. Or, inverting the order, they may also be called, respectively, chloride of sodium, sulphide of potassium, and oxide of barium.

Still another method, and one which has of late been very generally adopted, is to leave the name of the positive element unchanged, while giving the termination *ide* to the negative. By this method the examples just given would be sodium chloride, potassium sulphide, and barium oxide.

When two distinct compounds exist, consisting of the same two elements, as SO_2, SO_3, that one which contains the greater proportion of the negative element is named in accordance with the first method as given above, while the other simply changes the termination *ic* of the positive element to *ous*. SO_2 would then be sulphurous oxide and SO_3 sulphuric oxide.

When a compound exists which contains still less of the negative constituent than that indicated by the termination *ous*, as above, it receives the prefix *hypo;* as, N_2O_3 being nitrous oxide, N_2O is called hyponitrous oxide. Similarly the prefix *per* or hyper is used to indicate that the compound contains more of the negative element than the one whose positive constituent terminates in *ic*. This prefix is, however, usually given to the negative element; as, Mn_2O_3 being manganic oxide, MnO_2 is called manganic peroxide.

The prefix *sesqui* is sometimes used, as in Fe_2O_3, the sesqui oxide of iron.

The relative positions of the several compounds of a series may be illustrated by the following schedule, in which the plus sign indicates the root of the positive element and the minus sign that of the negative:

hypo + ous —— ide, as XO;
\+ ous —— ide, as XO_2;
hypo + ic —— ide, as XO_3;
\+ ic —— ide, as XO_4;
\+ ic per —— ide, as XO_5.

Unfortunately for the rule, however, the practice of recent authors varies very widely, so much so that it is the exception, for example, to find the whole series of oxides of nitrogen similarly named in any two works on chemistry.

In accordance with the above schedule the following would be the names of the five members of that series:

N_2O, hyponitrous oxide;
NO, nitrous oxide;
N_2O_3, hyponitric oxide;
NO_2, nitric oxide;
N_2O_5, nitric peroxide.

Instead of these, however, each member of the series has several different names in common use. For example, the first compound is known commonly as nitrous oxide; it is also called nitrogen monoxide. The second is variously termed nitric oxide, nitrosyl, nitrogen dioxide, since the molecule is by some considered to be indicated by the formula N_2O_2, and so on through the whole of the series.

Another striking example of exception to the rule will be found in the two compounds of carbon and oxygen, CO and CO_2. Clearly the former should be called carbonous oxide, and the latter carbonic oxide. Formerly, however, CO_2 was considered to be an acid, and was called carbonic acid; and in fact it is still so termed in most metallurgical and technical works, while CO was known as carbonic oxide. Since the adoption of the present views of the constitution of acids chemists no longer call CO_2 carbonic acid, and it is therefore usually named carbonic dioxide, or carbonic anhydride, from the fact that it may be considered as carbonic acid from which a molecule of water has been removed ($H_2CO_3 - H_2O = CO_2$). The name carbonic oxide would be inapplicable, since it would not be perfectly clear whether the compound indicated was CO or CO_2. As to CO, it still retains its old name carbonic oxide in some works, while in others it is termed carbonous oxide, or again carbon monoxide.

Still another source of difficulty to beginners exists in the fact that many compounds are known by irregular names; as, H_2O, water; NaCl, common salt; NH_3, ammonia, etc.

For the sake of uniformity it would perhaps be well if the system as shown in the following examples could be adopted:

N_2O, nitrogen monoxide;
N_2O_2, nitrogen dioxide;
N_2O_3, nitrogen trioxide;
N_2O_4, nitrogen tetroxide;
N_2O_5, nitrogen pentoxide;

or, nitric monoxide, etc.

Mn_3O_4 would, according to this plan, be called trimanganese tetroxide.

II. TERNARY COMPOUNDS.

a. ACIDS.—The elements hydrogen and oxygen are common to all of this class of acids. They therefore receive their specific names from the third or characteristic element; as, HNO_3 is called nitric acid. In case two or more acids occur, consisting of the same three elements united in different proportions, the prefixes and terminations, as explained under binary compounds, are applied, and in a similar manner; as,

HNO_2, nitrous acid;
HNO_3, nitric acid;
H_2SO_3, sulphurous acid;
H_2SO_4, sulphuric acid;
$HClO$, hypochlorous acid;
$HClO_2$, chlorous acid;
$HClO_3$, chloric acid;
$HClO_4$, perchloric acid.

b. SALTS.—The names of salts express both the metal contained and the acid from which they are derived. The metal takes the termination *ous* or *ic*. The termination of the acid is changed from *ous* to *ite*, and from *ic* to *ate;* as,

KNO_2, potassic nitrite;
$NaNO_3$, sodic nitrate;
$BaSO_4$, baric sulphate;
$FeSO_4$, ferrous sulphate;
$Fe_2(SO_4)_3$, ferric sulphate;
$KClO$, potassic hypochlorite;
$NaClO_2$, sodic chlorite;
$KClO_3$, potassic chlorate;
$KClO_4$, potassic perchlorate.

c. METALLIC HYDRATES.—The metal generally takes the termination *ic*, as NaHO, sodic hydrate. Where two hydrates of the same metal exist the one which has the lesser relative amount of the negative radical takes the termination *ous;* as, $Fe(HO)_2$, ferrous hydrate ; $Fe_2(HO)_6$, ferric hydrate.

In cases where only one hydrate is formed some authors prefer to retain the name of the metal unchanged ; as, sodium hydrate.

CLASSIFICATION OF ACIDS.

The hydrogen of an acid which is replaceable by a metal is called basic hydrogen.

Those acids which contain in the molecule but one atom of basic hydrogen are said to be monobasic ; those which contain two such atoms are dibasic ; those containing three are tribasic, and those containing four, tetrabasic. In hypophosphorous acid, H_3PO_2, we have an example of a monobasic acid, since only one of the three atoms of hydrogen is replaceable. In phosphorous acid, H_3PO_3, we have a dibasic acid, and in phosphoric acid, H_3PO_4, we have a tribasic acid, all three atoms of hydrogen being basic. Pyrophosphoric acid, $H_4P_2O_7$, is an example of a tetrabasic acid.

CLASSIFICATION OF SALTS.

Salts are either normal, acid, double, or basic.

A normal salt is one in which the whole of the basic hydrogen is replaced by the metal ; as, $NaNO_3$; Na_2SO_4 ; NaH_2PO_2 ; Na_2HPO_3 ; Na_3PO_4.

An acid salt is one in which some of the basic hydrogen still remains in the molecule ; as, $NaHSO_4$; NaH_2PO_3 ; Na_2HPO_4. There can of course be no acid salt of a monobasic acid, since, as the molecule contains but one atom of hydrogen, if any portion of that element is removed the whole must be.

A double salt is a complex molecule in which either the basic hydrogen is replaced by two distinct metals, or else the same metal replaces the basic hydrogen in two different acids. As examples of the first variety of double salt may be mentioned $(KCl)_2PtCl_4$, in which the two metals, potassium and platinum, replace the hydrogen in hydrochloric acid; $CaMg(CO_3)_2$; $K_2Al_2(SO_4)_4$. Examples of the second form,

$$PbSO_4, 3PbCO_3; PbCl_2, PbCO_3.$$

Basic salts are complex molecules which may be considered as consisting of a normal salt united to the oxide or the hydrate of the same metal; as, $BiCl_3$, Bi_2O_3; $Fe_2(SO_4)_3$, $Fe_2(HO)_6$; $(CuCO_3)_2$, $Cu(HO)_2$.

It will be seen by the foregoing examples how much more fully the formulas express the composition of bodies than the names.

The following list of chemical compounds contains the names and formulas of the substances most frequently used for testing in the laboratory. These names and formulas should be committed to memory, and the rules of nomenclature may be studied in their application to them:

Chlorhydric, or hydrochloric acid, HCl;
Nitric acid, HNO_3;
Sulphuric acid, H_2SO_4;
Sulphydric acid, H_2S;
Acetic acid, $HC_2H_3O_2$;
Ammonic hydrate, NH_4HO;
Ammonic sulphide, $(NH_4)_2S$;
Ammonic carbonate, $(NH_4)_2CO_3$;
Ammonic chloride, NH_4Cl;
Ammonic oxalate, $(NH_4)_2C_2O_4$;
Potassic ferrocyanide, $K_4(FeCy_6)$, or $K_4Fe(CN)_6$;
Potassic ferricyanide, $K_6(Fe_2Cy_{12})$, or $K_6Fe_2(CN)_{12}$;
Potassic sulphocyanate, $K(CyS)$, or $KCNS$;
Calcic hydrate, $Ca(HO)_2$;
Calcic chloride, $CaCl_2$;
Calcic sulphate, $CaSO_4$;
Baric chloride, $BaCl_2$;
Baric nitrate, $Ba(NO_3)_2$;
Baric carbonate, $BaCO_3$;
Magnesic sulphate, $MgSO_4$;

Ammonic molybdate, $(NH_4)_2MoO_4$;
Sodic hydrate, $NaHO$;
Sodic carbonate, Na_2CO_3;
Disodic hydric phosphate, Na_2HPO_4;
Sodic acetate, $NaC_2H_3O_2$;
Potassic dichromate, $K_2Cr_2O_7$;
Ferrous sulphate, $FeSO_4$;
Ferric chloride, Fe_2Cl_6;
Cobaltic nitrate, $Co(NO_3)_2$;
Plumbic acetate; $Pb(C_2H_3O_2)_2$;
Argentic nitrate, $AgNO_3$;
Mercuric chloride, $HgCl_2$;
Platinic chloride, $PtCl_4$;
Stannous chloride, $SnCl_2$;
Alcohol, C_2H_6O.

After having made himself familiar with the formulas of the substances with which he tests, the student should write in the form of an equation the result of the action of each test which he performs upon a compound under examination.

CHAPTER III.

CHEMICAL OPERATIONS.

Reaction with Test Paper. A small piece of red or blue litmus paper is dipped in the solution. If it is acid, blue paper is turned red; if it is alkaline, red paper is turned blue.

Turmeric paper is turned brown by alkalies.

Precipitation. When an insoluble body is formed in a solution and separates (falls) from it, precipitation is said to take place. Precipitates are *gelatinous*, as aluminic hydrate; *flocculent* (consisting of flakes), as sulphide of zinc; or *pulverulent*, as baric sulphate. Usually the particles are less finely divided, and filtration is easier with precipitates which form in dilute solutions, particularly in boiling solutions. With some precipitates, as magnesic phosphate, in very dilute solutions, the act of precipitation is a slow process of crystallization, and the formation of a precipitate does not take place until after several hours.

Filtration is a process by which an insoluble body, usually a precipitate, is separated from a liquid. It is usually important to allow a precipitate to settle before filtration; and frequently after the clear liquid has been poured upon the filter it is best to add more water to the precipitate, and to wait again until it has settled. When the precipitate requires to be washed, this process may be repeated many times before the precipitate is brought upon the filter. The precipitate usually clogs the pores of the filter, so that it retards the flow of the liquid. A precipitate which has to be washed is finally brought on the filter by rinsing with the wash-bottle, and it is washed by repeatedly filling the filter with water and allowing it to empty itself. The process should be intermittent, and the filter should never be kept constantly full during the latter part of a filtration, when the object is to wash a precipitate with pure water. Sometimes precipitates which take the form of a powder are so finely divided as to pass through the pores of a filter. This can usually be avoided by precipitating in a dilute solution, and particularly by boiling the solution. Sulphur cannot be prevented from going through the filter. Filtration is more rapid with hot than with cold water. Before commencing a filtration, the filter should be made to fit closely to the side of the funnel, and it should always be moistened with pure water.

Decantation consists in allowing a precipitate to settle and in pouring off the liquid above it, in the manner already mentioned under filtration. This process may or may not be united with that of filtration. For instance, argentic chloride settles so completely and quickly that it can usually be washed simply by repeated decantations.

Evaporation is usually performed in a porcelain dish. A few drops of liquid can be evaporated by heating them on platinum foil. Bits of broken glass or porcelain vessels are very useful for the same purpose.

The Use of the Blowpipe. An olive oil or kerosene lamp, with a wick $\frac{5}{8}$ in. long and $\frac{1}{4}$ in. broad, or a Bunsen's lamp, with the regulator turned to shut off the draught of air, may be used. If the Bunsen's lamp has no regulator for the draught, a smaller tube may be introduced into the lamp-tube, until it rests upon the piece from which the gas issues and excludes the air. After considerable practice a continuous stream of air can be forced through the blowpipe by making the cavity of the mouth the reservoir, into which air is forced at intervals from the lungs, and is prevented from escaping by a peculiar contraction of the throat and hanging palate, which is easily learned ; the breathing goes on through the nose uninterruptedly. The chief difficulty that beginners usually experience is fatigue of the muscles of the cheeks, which prevents a long-continued effort. Two kinds of flames can be produced with the blowpipe—one containing an excess of air, consequently of oxygen ; the other containing an excess of combustible gases, consequently of gases capable of consuming oxygen or reducing.

The Oxidizing Flame is produced by introducing the point of the blowpipe one-third through the lamp-flame. It is a *clear blue* cone, surrounded and continued at the point by a colorless flame, intensely hot, and capable of producing oxidation. The substance should be heated beyond the point of the blue cone.

The Reducing Flame is produced by holding the point of the blowpipe at the outside of the lamp-flame, and by blowing somewhat more gently. The flame is much less pointed and is more luminous than the oxidizing flame. The substance should be heated at a distance from the point of the flame equal to one-third of its length, and should be completely enveloped in the flame. The position of the blowpipe and the force of the blast regulate the quality of the flame. It is a difficult matter to produce the true reducing flame, which should not deposit carbon on the substance heated in it, and at the same time should contain no excess of oxygen.

The Borax Bead is formed by making a loop $\frac{1}{8}$ in. in diameter in a piece of platinum wire, heating it red hot in the blowpipe flame, and touching it to a small piece of borax while it is hot. The borax, which adheres to the hot wire, is heated in the blowpipe flame. It at first swells while losing its water of crystallization, and finally it melts to a clear glass bead. A finely divided substance can be taken up by touching the hot bead to it, and it can then be tested as to its solubility in the borax bead, coloring properties, etc., in the blowpipe flame.

Heating on Charcoal. Select a good piece of charcoal, at least 4 in. long and 1 in. broad and thick, and smooth a plane surface in a direction at a right angle with that of the year-rings. (If heat is applied to a surface parallel to the planes of the year-rings, the charcoal is more liable to snap from the expulsion of moisture.) In many cases the charcoal serves as a convenient support for a substance to be heated; in others the reducing agency of the charcoal comes in play. Substances are also evaporated at a high temperature from the surface of the charcoal.

Ductility, Malleability, Brittleness are characteristic properties of metals, and metals can be tested with regard to them by pounding with a hammer or by rubbing with the pestle of a mortar. When a substance which appears to contain a higher metal is reduced by sodic carbonate on charcoal, unless metallic globules are at once apparent, the portion of the charcoal which has been heated should be cut out and pulverized with water in a mortar, and washed by decantation. If metallic globules have been formed, they will sink to the bottom, and after thorough washing, during which the sodic carbonate is dissolved, and the light particles of charcoal are floated away, they will appear as globules, if they are hard like copper; as brittle grains, if they are brittle like bismuth; or as flattened disks, when they are ductile like lead, and when they have been pressed by the pestle against the mortar.

Color of the Flame. If the substance to be tested is a solid, a small piece of it is brought on a loop of fine platinum wire, or in a pair of forceps, into the flame of the alcohol or Bunsen lamp, and the color imparted to the flame observed. If a substance in solution is to be tested, the platinum wire is dipped in the solution and is then introduced into the lamp-flame. If the solution is too dilute to afford a distinct test in this way, it must be evaporated, and it is usually best to evaporate nearly to dryness, and to take some of the solid residue for the test.

The Manipulation of Glass Tubing. Glass softens when a small piece is heated in the flame of an alcohol lamp, or when a larger piece is heated in a Bunsen lamp, or with the blowpipe flame or in a blast lamp.

A tube can be bent easily as soon as the glass softens.

It is best only to bend gently at first, then to heat the adjacent part of the tube, and to bend again, and so on, in order that the sides of the tube may not fall together in bending.

When a glass tube is heated for some time, it contracts and the sides thicken. By drawing out a tube either immediately after it has become soft, or after the sides have thickened, a tapering point of any desired calibre and thickness of glass can be obtained.

To close the end of a glass tube, draw the tube off while the glass is as thin as possible, and hold the tapering point in the flame, and draw the end off again ; in this way a tube with a pointed closed end is obtained ; by heating the closed end of the tube, removing it from the flame, rotating it and blowing in it, while the glass is still red hot, it expands, and a more rounded end or a bulb can be produced. Glass tubing can be cut by making a mark at the required place by a few file strokes, and then by breaking the glass. The ends of glass tubes cut in this way should be held in the flame till they become red hot ; in this way the sharp edges become rounded.

PART II.

PART II. is preparatory to Part III., which contains a general scheme of analysis applicable to compounds of all the elements. The most important tests are those which are described in Part III., and a knowledge of them would suffice alone for the purposes of analysis, if the liability to error in chemical manipulations did not make it expedient to employ a variety of tests, as corroborative evidence, before coming to a conclusion in regard to the composition of a substance.

It is important that the student should turn to Part III., and commit to memory the general features of the scheme of separation for each group, at the time that he is performing the reactions of the members of the group as they are described in the following pages. By this means he will make himself familiar with the important points in which the compounds with which he has to deal differ from each other, and the manner in which these differences can be used in analysis ; also at this stage of his progress it is advisable for him to make mixtures of compounds of several elements, and to analyze them according to the directions given in Part III.

The student, keeping in view the reasons for learning the characteristic reactions of the compounds of each element, should perform carefully the tests described in Part II., supplementing the description by the closest observation of the phenomena as they pass before his eyes. By practice of this kind he will soon acquire the skill in manipulation necessary for analytical work. ALWAYS, WHEN A REACTION IS PERFORMED, THE EQUATION DESCRIBING IT SHOULD BE WRITTEN. The formulas of the reagents and the compound operated upon* form the first half of the equation; the formula of the precipitate, which is given in the book, enters into the second half of the equation and determines the formulas of its other members.

Thus it is known that baric chloride and calcic sulphate give a precipitate of BARIC SULPHATE, $BaSO_4$. (See page 34.) From the formulas on the labels of the bottles we can construct the equation: $BaCl_2 + CaSO_4 = BaSO_4 + X$, and by inspecting the equation we find the unknown quantity X can only be $CaCl_2$. The following case is more complicated: disodic hydric phosphate, ammonic hydrate, and magnesic sulphate form a precipitate of MAGNESIC AMMONIC PHOSPHATE, $MgNH_4PO_4$; or, putting the statement into formulas, $Na_2HPO_4 + NH_4HO + MgSO_4 = MgNH_4PO_4 + X$. Here $X = Na_2 + H_2 + O + SO_4$, and the question arises: How are these bodies combined? A slight experience will teach that the rule 2d (page 9) brings the SO_4 and the Na together, and consequently the H_2 and the O; while an inspection of the tables of quantivalence (pages 2 and 12) shows that SO_4 combines with 2Na, and that O combines with 2H; hence X becomes $Na_2SO_4 + H_2O$.

The grouping together of the elements or groups of elements appearing in reactions is not usually a difficult matter, and is soon learned with practice.

The formulas of the compounds which are most frequently used in the laboratory stand after the names of the metals and acids, and can be used in writing equations.

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* These formulas should be given in full upon the labels of the bottles containing the compounds used and the reagents.

TESTS FOR METALS.

GROUP I.

SODIUM, POTASSIUM, AND AMMONIUM.

THERE is no reagent which precipitates all the metals of this group. The salts of metals of Group I. have a neutral reaction when they contain strong acids like chlorhydric, nitric, and sulphuric acids. They have an alkaline reaction when they contain weak acids like sulphydric, boracic, and carbonic acids.

SODIUM.

$NaCl$; Na_2CO_3; $NaHO$.

Sodium compounds can be recognized by heating them in the loop of a piece of fine platinum wire in the flame of a lamp. Sodium colors the flame *yellow*, and can be recognized, even when mixed with much larger quantities of other elements, which alone impart other colors to the flame.

When a liquid is to be tested, it may be evaporated and the residue brought on the platinum wire, or frequently it is sufficient to dip the wire in the liquid and to bring it into the flame of the lamp.

No reagent * precipitates sodium compounds.

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* Only the reagents spoken of in this book are referred to.

POTASSIUM.

KCl; K_2SO_4.

Potassium compounds color the flame of a lamp *violet*. A small admixture of sodium obscures the color of the flame of potassium, but the sodium color disappears, and that characteristic of potassium can be observed when the flame is viewed through a thick glass colored blue with cobalt.

Mix together a sodium and a potassium salt, observe the yellow color imparted to the flame by the mixture, showing the presence of sodium, and then examine the flame for the potassium color through a piece of cobalt glass thick enough to exclude the sodium flame. Examine a pure sodium flame with cobalt glass to ascertain that the glass does not allow the color of the sodium to pass through it, or until the blue color of the sodium flame can be easily distinguished from the violet of the potassium flame.*

Platinic Chloride precipitates concentrated solutions of potassic chloride as a DOUBLE CHLORIDE OF PLATINUM and POTASSIUM, $(KCl)_2 PtCl_4$. No other salt of potassium can be used for this test. It is best to evaporate the solution to dryness in a water-bath, with a large quantity of platinic chloride, and to wash the residue several times with alcohol. The double chloride is left as a *yellow crystalline powder*, which gives the potassium flame. The double chloride is slightly soluble in water, but insoluble in alcohol.

No other reagent precipitates potassium compounds.

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* If the glass is sufficiently thick and intense in color, the light from the sodium flame is completely excluded. A thinner glass allows a part of the light to pass through, but it then has a blue color which can be distinguished after practice from the violet color of the potassium flame, which passes through the glass with little of its brilliancy diminished.

AMMONIUM.

NH_4Cl; NH_4HO.

Ammonium compounds do not color the flame of a lamp.

Platinic Chloride precipitates ammonic chloride as a DOUBLE CHLORIDE OF PLATINUM AND AMMONIUM $(NH_4Cl)_2 PtCl_4$. The precipitate forms under the same circumstances, and has the same aspect and properties as that obtained with potassic chloride, but it can be distinguished from the latter by the absence of color imparted to the flame of a lamp when it is heated in it. It is destroyed at a dull red heat, and a residue of metallic platinum is left.

Sodic Hydrate, added in excess to ammonic compounds, causes them to give off the smell of AMMONIA GAS, NH_3, especially when the solution is heated. AMMONIA GAS colors moist turmeric paper *brown* (a delicate test).

No other compound interferes with the application of this test.

GROUP II.

BARIUM, STRONTIUM, CALCIUM, AND MAGNESIUM.

THE chlorides and nitrates of metals of Group II. are soluble in water. The sulphates of calcium and magnesium are also soluble. The solutions have a neutral reaction with test-paper.

Ammonic and Sodic Carbonates precipitate the metals of Group II. in neutral solutions as CARBONATES. The carbonate of magnesium is very soluble in solutions of ammonic salts, particularly in ammonic chloride ; therefore no precipitate of magnesic carbonate is produced when these salts are present in considerable quantity.

Neutral Phosphates (as disodic hydric phosphate) precipitate the metals of Group II. in neutral or alkaline solutions as PHOSPHATES.

The carbonates and phosphates of metals of Group II. are soluble in dilute acids, unless the acids themselves are capable of precipitating the metals.

Sulphydric Acid, Ammonic Sulphide, and Ammonic Hydrate do not precipitate the metals of Group II.

BARIUM.

$BaCl_2$; $Ba(NO_3)_2$; $Ba(HO)_2$.

Sulphuric Acid (dilute) precipitates baric compounds, as BARIC SULPHATE, $BaSO_4$, *white powder.*

Calcic Sulphate and other soluble sulphates give the same precipitate with baric compounds.

Baric sulphate is insoluble in acids.

Ammonic Oxalate precipitates baric compounds from concentrated neutral or alkaline solutions, as BARIC OXALATE, BaC_2O_4, *white powder*, soluble in acids.

Barium compounds, particularly when moistened with chlorhydric acid, color the flame *yellowish green.*

STRONTIUM.

$SrCl_2$; $Sr(NO_3)_2$.

Sulphuric Acid (dilute) and soluble sulphates precipitate strontium from its solutions, as STRONTIUM SULPHATE, $SrSO_4$, *white powder.* The precipitate will not appear immediately unless the solution be concentrated. In very dilute solutions it may not appear at all, since strontium sulphate is slightly soluble in water (1 part in 6900).

Ammonic Oxalate precipitates strontium compounds from neutral or alkaline solutions, as STRONTIUM OXALATE, SrC_2O_4, *white powder*, soluble in acids.

Strontium compounds, particularly when moistened with chlorhydric acid, color the flame *brilliant red.*

CALCIUM.

$CaCl_2$; $CaSO_4$; $Ca(HO)_2$.

Sulphuric Acid (dilute) and soluble sulphates, except calcic sulphate, precipitate concentrated solutions of calcium compounds, as CALCIC SULPHATE, $CaSO_4$, *white powder.* Calcic sulphate is soluble in a considerable quantity of water, therefore no precipitate is produced in very dilute solutions by sulphuric acid. It is insoluble in dilute alcohol, therefore sulphuric acid (dilute), with the addition of a large quantity of alcohol, precipitates calcic sulphate from even dilute solutions of calcium compounds.

Ammonic Oxalate precipitates calcic compounds from neutral or alkaline solutions, as CALCIC OXALATE, CaC_2O_4, *white powder.* The precipitate forms best after standing some time in a solution to which ammonic hydrate has been added in excess. Calcic oxalate does not dissolve in acetic acid.

Calcium compounds, particularly when moistened with chlorhydric acid, color the flame *yellowish red.*

MAGNESIUM.

$MgCl_2$; $MgSO_4$.

Hydro-Disodic Phosphate precipitates magnesic compounds, to whose solution ammonic hydrate and ammonic chloride have been added, as MAGNESIC AMMONIC PHOSPHATE, $MgNH_4PO_4$, *white crystalline powder, or white flakes if the solution is concentrated.* This precipitate only appears after the lapse of some time in very dilute solutions. It is then crystalline.

Sodic Hydrate, in excess, precipitates magnesic compounds, as MAGNESIC HYDRATE, $Mg(HO)_2$, *white powder*, when it is boiled with their solutions. In case ammonic chloride is present, the boiling must be continued until the odor of ammonia is no longer perceptible.

Mix together the chlorides of barium, strontium, calcium and magnesium. Add a small quantity of ammonic chloride, then ammonic hydrate, and finally ammonic carbonate. The barium, strontium, and calcium will be precipitated as carbonates. Filter, and to the filtrate add hydro-disodic phosphate. The magnesium will be precipitated as magnesic ammonic phosphate. Wash the precipitate, consisting of the carbonates of barium, strontium, and calcium, with water, and pour over it a small quantity of dilute hydrochloric acid. This will decompose the carbonates, and the chlorides of the three metals will be formed, which being soluble will pass through the filter and can be caught in a test-tube or small beaker. Dilute the solution with considerable water and add dilute sulphuric acid. Barium and strontium will be precipitated as sulphates, while the calcic sulphate will remain in solution. Filter, add ammonic hydrate to the filtrate until the reaction becomes alkaline, and then ammonic oxalate. The calcium will be precipitated as oxalate. Wash the sulphates of barium and strontium on the filter, and then moisten a small particle of the mass with chlorhydric acid and examine it on platinum wire in the flame. The yellowish-green color of barium will first appear. After holding the wire in the flame for a few seconds dip it into chlorhydric acid and again bring it into the flame. By repeating this operation several times the barium flame will disappear and the crimson color, characteristic of strontium, will be very plainly visible.*

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* Processes similar to the above are used for the separation of all the metals from each other, and care must always be taken to ascertain whether enough of a reagent has been added to completely effect a precipitation before the next test is proceeded with in the filtrate, and when a precipitate is to be examined, it must be thoroughly freed by washing from the solution which adheres to it.

GROUP III.

ALUMINIUM AND CHROMIUM.

THE sulphates, chlorides, and nitrates of metals of Group III. are soluble in water, and the solution has an acid reaction with test-paper. Aluminic and chromic alum solutions have a neutral reaction.

Ammonic Hydrate, Carbonate, and Sulphide precipitate the metals of Group III. as HYDRATES.

Neutral Phosphates precipitate the metals of Group III. as PHOSPHATES.

Sulphydric Acid does not precipitate the metals of Group III.

ALUMINIUM.

$Al_2(SO_4)_3$; Al_2Cl_6.

Ammonic Hydrate precipitates aluminic compounds as ALUMINIC HYDRATE, $Al_2(HO)_6$, *gelatinous, white flakes.*

The precipitate is insoluble in ammonic hydrate and in ammonic chloride, but dissolves in acids and in sodic hydrate. It forms best on boiling.

Sodic Hydrate precipitates aluminic compounds like ammonic hydrate, but an excess of sodic hydrate dissolves the precipitate so quickly that its formation easily escapes notice.

No precipitate is formed when the solution in sodic hydrate is boiled.

Solid compounds of aluminium (except silicates), when moistened with cobaltic nitrate solution, and heated with the oxidizing blowpipe flame (see page 25) on charcoal, or on a platinum wire, take a *blue color.*

CHROMIUM.

$Cr_2(SO_4)_3$; Cr_2Cl_6.

Solutions of chromic oxide compounds are *green.*

Ammonic Hydrate precipitates chromic compounds as CHROMIC HYDRATE, $Cr_2(HO)_6$, *gelatinous, dirty green flakes.* The precipitate is insoluble in ammonic hydrate after boiling, and in ammonic chloride, but dissolves in acids and in sodic hydrate.

Sodic Hydrate precipitates chromic compounds like ammonic hydrate, but dissolves them when an excess of sodic hydrate is present. CHROMIC HYDRATE is precipitated from its solution in sodic hydrate when the *dilute* solution is boiled for some time.

Blowpipe Reactions. Compounds of chromium color the borax bead *green.*

If chromic hydrate, or any solid chromic compound, is mixed with equal parts of sodic carbonate and sodic nitrate, and heated red hot on the platinum foil, chromate of sodium is formed by the oxidation of chromic oxide. Chromate of sodium dissolves in water with a yellow color. The color is heightened when an acid is added, and an acid chromate is formed in the solution.

This reaction is a characteristic test for chromium compounds.

Mix together solutions of chromium and aluminium salts, add sodic hydrate until the reaction becomes very strongly alkaline (the precipitate which first forms will dissolve), dilute with a considerable quantity of water in a small flask, and boil for several minutes after a dirty green precipitate has formed. Filter from the precipitate. This precipitate contains all the *chromium.* Test it according to Part III. (***102***). The filtrate contains all the *aluminium.* Test it according to Part III. (***104***).

GROUP IV.

ZINC, MANGANESE, IRON, NICKEL, AND COBALT.

THE sulphates, chlorides, and nitrates of metals of Group IV. are soluble in water. The solutions have an acid reaction with test-paper.

Ammonic Sulphide precipitates the metals of Group IV. as sulphides ; if the solution is not neutral, it should be made so with ammonic hydrate.

Sodic Hydrate and Ammonic Hydrate precipitate the metals of Group IV. as HYDRATES. The HYDRATE OF ZINC is soluble in an excess of the precipitant, and the HYDRATES OF NICKEL AND COBALT are soluble in ammonic hydrate.

Sodic Carbonate precipitates the metals of Group IV. as CARBONATES (ferric compounds as ferric hydrate).

Neutral Phosphates precipitate the metals of Group IV. as PHOSPHATES.

Sulphydric Acid does not precipitate the metals of Group IV. when they are in an acid solution. (See Zinc, page 40.)

SECTION I.

ZINC, MANGANESE, AND IRON.

Metals whose sulphides are soluble in cold dilute chlorhydric acid.

ZINC.

$ZnSO_4$; $ZnCl_2$.

Metallic zinc dissolves readily in dilute chlorhydric acid and sulphuric acid, with evolution of hydrogen.

The metal melts readily when heated on charcoal with the blowpipe, and at a high temperature it distils, and the vapor burns with a *bluish-white* flame, depositing an incrustation on the charcoal of OXIDE OF ZINC, ZnO, which is *white* when hot and *yellow* when cold. If the incrustation is moistened with cobaltic nitrate and heated in the oxidizing flame, it turns *dirty green.* By this test zinc can often be recognized in alloys.

Ammonic Sulphide precipitates zinc compounds as the SULPHIDE OF ZINC, ZnS, *white, flocculent precipitate.* The sulphide of zinc is soluble in dilute chlorhydric acid, but not in acetic acid. It is the only white insoluble sulphide.

Sulphydric Acid precipitates zinc as the SULPHIDE OF ZINC only when the metal is combined with acetic acid. To obtain the precipitate, if a stronger acid is present, add sodic hydrate until the solution has a strongly alkaline reaction, and then add acetic acid, until the reaction becomes acid, before treating with sulphydric acid.

Sodic and Ammonic Hydrates precipitate zinc compounds as the HYDRATE OF ZINC, $Zn(HO)_2$, *white flakes.* The precipitate dissolves easily in an excess of the precipitants ; and the solution in sodic hydrate is not reprecipitated when it is boiled.

MANGANESE.

$MnSO_4$; $MnCl_2$.

Solutions of manganese compounds have a *faint pink* color.

Ammonic Sulphide precipitates manganese compounds as the SULPHIDE OF MANGANESE, MnS, *flesh-colored flakes.* The sulphide of manganese is soluble in dilute acids.

Sodic and Ammonic Hydrates precipitate manganese compounds as MANGANOUS HYDRATE, $Mn(HO)_2$, *white flakes, which turn brown on exposure to the air.* Manganous hydrate is insoluble in an excess of the precipitant, but it is soluble in a large quantity of ammonic chloride.

Blowpipe Reactions. Manganese compounds color the borax bead *amethyst* in the oxidizing flame. If a compound of manganese is heated with a soda bead (which can be made in the same way as a borax bead in the loop of a platinum wire) in the oxidizing flame, it colors it *green*, in consequence of the formation of manganate of sodium. The same color is produced when a compound of manganese is heated on the platinum foil with sodic carbonate and nitrate. On boiling the green salt with water containing a little alcohol it is destroyed, the color disappears, and brown flakes of manganic hydrate are precipitated.

IRON (*ferrous salts*).

$FeSO_4$; $FeCl_2$.

Metallic iron dissolves readily in dilute acids, with evolution of hydrogen.

Ferrous salts in solution have a *pale green* color.

Oxidizing Agents (as nitric acid and potassic chlorate), when heated with acid solutions of ferrous salts, oxidize them to FERRIC SALTS, whose color is *brownish red*, or *reddish yellow*, and is more intense than the *green* color of ferrous salts.

Ammonic Sulphide precipitates ferrous salts as FERROUS SULPHIDE, FeS, *black flakes.* Ferrous sulphide is soluble in dilute acids.

Sodic and Ammonic Hydrates precipitate ferrous compounds as FERROUS HYDRATE, $Fe(HO)_2$, *at first nearly white, then bluish green, and, finally, by absorption of oxygen, reddish brown.* Ferrous hydrate is insoluble in an excess of sodic hydrate. The presence of a large quantity of ammonic salts in a solution prevents its precipitation.

Potassic Ferrocyanide precipitates ferrous compounds as POTASSIC FERROUS FERROCYANIDE, $K_2Fe(FeCy_6)$,*

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* Cy, cyanogen, is used as a symbol for the group CN.

bluish white, turning quickly *dark blue*, through absorption of oxygen from the air.

Potassic Ferricyanide precipitates ferrous compounds, as TURNBULL'S BLUE, $Fe_3(Fe_2Cy_{12})$,* *deep blue. This is the best test for ferrous compounds.*

The last two precipitates are insoluble in dilute acids.

Potassic Sulphocyanate gives no coloration with ferrous compounds.

IRON (*ferric salts*).

Fe_2Cl_6.

Reducing Agents, as sulphurous and sulphydric acids, metallic zinc and iron, reduce ferric salts in solution to ferrous salts when a free acid is present.

The reaction with sulphydric acid is accompanied by a precipitation of sulphur, $Fe_2Cl_6 + H_2S = 2FeCl_2 + 2HCl + S$.

A similar reaction takes place with ammonic sulphide and the other salts of sulphydric acid.

Ferric salts in solution have a *yellow* color, and they possess a much stronger coloring power than ferrous salts.

Ammonic Sulphide reduces ferric salts to ferrous salts, and then precipitates FERROUS SULPHIDE, FeS, *black flakes*.

Sodic and Ammonic Hydrates precipitate ferric compounds as FERRIC HYDRATE, $Fe_2(HO)_6$, *red gelatinous flakes*, insoluble in an excess of the precipitants and in ammonic salts.

Potassic Ferrocyanide produces a precipitate of PRUSSIAN BLUE, $Fe_4(FeCy_6)_3$, *deep blue*, in a solution of ferric salts.

Potassic Ferricyanide colors solutions of ferric salts *deep reddish brown*, but produces no precipitate. On the addition of a reducing agent a *deep blue* precipitate forms.

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* See note on preceding page.

Potassic Sulphocyanate gives a *deep red color*,* with the smallest traces of ferric compounds in acid solutions.

The different behavior with the last three reagents of ferrous and ferric salts serves to distinguish between them.

Blowpipe Reaction. Iron colors the borax bead *green* in the reducing flame, and *reddish yellow while hot*, and *yellow while cold*, in the oxidizing flame.

Mix together solutions of zinc, manganese, and iron salts ; if a ferrous salt is taken, add a little chlorhydric acid and boil the solution for a few minutes with one or two crystals of potassic chlorate, in order to convert the ferrous into ferric salt. Add sodic hydrate to the solution until the reaction is strongly alkaline, boil for a few minutes and filter. All the *manganese* and *iron* will be precipitated. Test the precipitate for iron according to Part III. (***101***), and for manganese according to Part III. (***102***). All the *zinc* will be contained in the filtrate ; test it for zinc according to Part III. (***103***).

SECTION II.

NICKEL AND COBALT.

Metals whose sulphides are insoluble in cold, dilute chlorhydric acid.

NICKEL.

$NiSO_4$; $Ni(NO_3)_2$; $NiCl_2$.

Solutions of nickel salts are *green.*

Ammonic Sulphide precipitates nickel compounds as the SULPHIDE OF NICKEL, NiS, *black flakes.* The sulphide of nickel is insoluble in cold, dilute chlorhydric acid. It dissolves readily on boiling, or in a strong acid.

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* Potassic sulphocyanate gives the same color with a solution containing a large quantity of free nitric acid.

Sodic and Ammonic Hydrates precipitate nickel compounds as the HYDRATE OF NICKEL, $Ni(HO)_2$, *apple green.* The hydrate of nickel is insoluble in an excess of sodic hydrate. It dissolves in ammonic hydrate, and the solution has a *blue* color.

Blowpipe Reactions. Nickel colors the borax bead in the oxidizing flame *violet, when it is hot, and a faint reddish-brown when it is cold.* By long-continued reduction in the reducing flame, or on charcoal, the bead may be obtained *colorless*, but with gray specks of reduced metal in it.

COBALT.

$Co(NO_3)_2$; $CoCl_2$.

Solutions of cobalt salts, when dilute, are *red.*

Ammonic Sulphide precipitates cobalt compounds as the SULPHIDE OF COBALT, CoS, *black flakes.* The sulphide of cobalt is insoluble in cold dilute chlorhydric acid. It dissolves readily on boiling or in a strong acid.

Sodic and Ammonic Hydrates precipitate cobalt compounds at first as a *blue basic salt*, which changes to the *pale red* COBALTOUS HYDRATE, $Co(HO)_2$ on boiling. On exposure to the air it becomes brown, through the formation of cobaltic hydrate. Cobaltous hydrate is insoluble in an excess of sodic hydrate, but it dissolves in ammonic hydrate, and the solution is *red, tinged with brown.*

Blowpipe Reactions. Cobalt compounds color the borax bead *blue*, and the color is so intense that a small quantity of cobalt eclipses the color produced by a much larger quantity of nickel, when the latter is mixed with it. The blue color does not disappear on reduction, so that when sufficient nickel is present to hide the color of a small quantity of cobalt in a bead, the color of the nickel may be made to disappear by a thorough reduction, either on the platinum wire or on charcoal, so that the blue color characteristic of cobalt can be detected in the bead. If the bead was reduced on charcoal, it is advisable to remove it from the charcoal and to melt it on the platinum wire in order to examine its color.

Mix together sulphates of zinc, manganese, and nickel, and the nitrate of cobalt and ferrous sulphate, add ammonic hydrate until a permanent precipitate begins to form, and then ammonic sulphide * until the metals are completely precipitated as sulphides. Wash the precipitate on a filter and treat it with cold, dilute chlorhydric acid, in order to separate the sulphides of *nickel* and *cobalt* from the other sulphides. See Part III. (***96***).

Test for nickel and cobalt according to Part III. (***97***) and (***98***).

Test for the other metals in the chlorhydric acid solution according to Part III. (***99***), (***100***), (***101***), (***102***), and (***103***).

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* In order to precipitate the nickel completely, the solution must not contain free ammonic hydrate, and the ammonic sulphide must not contain an excess of ammonic hydrate. If these precautions are not observed a part of the sulphide of nickel dissolves, imparting a brown color to the solution.

GROUP V.

SILVER, MERCURY, LEAD, COPPER, BISMUTH, AND CADMIUM.

THE salts of metals of Group V. with chlorhydric, nitric, and sulphuric acids, which are soluble in water, have an acid reaction with test-paper.

Sulphydric Acid and Ammonic Sulphide precipitate metals of Group V. in neutral or acid solutions as SULPHIDES. The sulphides of metals of Group V. are insoluble in AMMONIC SULPHIDE and in SODIC HYDRATE, and in dilute acids, even when heated ; but, with the exception of mercuric sulphide, they are all dissolved when boiled with strong NITRIC ACID.

Sodic and Ammonic Hydrates precipitate metals of Group V. as OXIDES or HYDRATES. The hydrate of lead is somewhat soluble in an excess of sodic hydrate, and the oxide of silver and the hydrates of copper and cadmium are readily soluble in an excess of ammonic hydrate.

Sodic Carbonate precipitates the metals of Group V. as CARBONATES.

Neutral Phosphates precipitate metals of Group V. in neutral solutions as PHOSPHATES.

SECTION I.

SILVER, MERCUROUS COMPOUNDS, AND LEAD.

Metals whose compounds are precipitated as chlorides by chlorhydric acid.

SILVER.

$AgNO_3$.

Metallic silver and its alloy with copper dissolve readily in nitric acid.

Sulphydric Acid and Ammonic Sulphide precipitate compounds of silver as the SULPHIDE OF SILVER, Ag_2S, *black flakes.*

Chlorhydric Acid precipitates compounds of silver as the CHLORIDE OF SILVER, $AgCl$, *white flakes, which settle readily after boiling or prolonged shaking.* The chloride of silver dissolves readily in ammonic hydrate. It is insoluble in concentrated nitric acid, even when it is boiled with it.

Sodic and Ammonic Hydrates precipitate compounds of silver as the OXIDE OF SILVER, Ag_2O, *a grayish brown powder.*

The oxide of silver dissolves very readily in ammonic hydrate.

Blowpipe Reactions. Compounds of silver, mixed with sodic carbonate, are easily reduced on charcoal, by heating in the blowpipe flame, and a hard globule of metallic silver is obtained.

MERCUROUS COMPOUNDS.

$Hg_2(NO_3)_2$; Hg_2Cl_2.

Metallic mercury dissolves readily in nitric acid to form mercurous nitrate.

Copper in the form of a thin sheet or wire can be coated with mercury by immersing it in a solution of mercury containing a free acid. The coating is deposited in a longer or shorter time, according to the strength of the solution. It takes the color of mercury when it is rubbed gently with a bit of cloth or paper. The copper should be cleansed by immersing it in dilute nitric acid solution before it is put in the solution containing mercury.

Oxidizing Agents, as chlorine, aqua regia, or strong nitric acid, transform mercurous into mercuric compounds.

Sulphydric Acid and Ammonic Sulphide precipitate mercurous compounds as MERCUROUS SULPHIDE, Hg_2S, *black*. Mercurous sulphide is not dissolved when it is boiled with moderately strong nitric acid. It dissolves readily in aqua regia.

Chlorhydric Acid precipitates mercurous compounds as MERCUROUS CHLORIDE, Hg_2Cl_2, *white powder*. Mercurous chloride turns black, but does not dissolve when ammonic hydrate is added to it. It is insoluble in dilute acids. It dissolves in aqua regia.

Sodic and Ammonic Hydrates give with mercurous compounds *black* precipitates, insoluble in an excess of the precipitant. (For tests by heating mercurous compounds, see page 50.)

LEAD.

$$Pb(C_2H_3O_2)_2\ ;\ Pb(NO_3)_2.$$

Metallic lead dissolves readily in nitric acid.

Sulphydric Acid and Ammonic Sulphide precipitate lead compounds as the SULPHIDE OF LEAD, PbS, *black*. Sulphide of lead is oxidized by strong nitric acid, with formation of sulphate of lead, *white powder*, which is insoluble, unless a very large quantity of nitric acid is present.

Chlorhydric Acid precipitates lead compounds as the CHLORIDE OF LEAD, $PbCl_2$, *white*. When the solution is dilute no precipitate is produced. The chloride of lead dissolves entirely on boiling with a large quantity of water.

Sulphuric Acid precipitates lead compounds as the SULPHATE OF LEAD, $PbSO_4$, *white powder*. Sulphate of lead is soluble to some extent in chlorhydric and nitric acids, and it is slightly soluble in water. It is insoluble in a mixture of alcohol and water. When complete precipitation is required, it is best to evaporate with an excess of sulphuric acid until all the other acids are driven off, then to dilute with water, and to add an equal bulk of alcohol.

LEAD, BARIUM, STRONTIUM, CALCIUM, *and* STANNIC COMPOUNDS *are the only ones precipitated by* SULPHURIC ACID.

Sodic and Ammonic Hydrates precipitate compounds of lead as BASIC SALTS OF LEAD, *white precipitate.* Basic salts of lead are somewhat soluble in sodic hydrate.

Blowpipe Reactions. Solid compounds of lead give, when heated with sodic carbonate on charcoal, globules of metallic lead, recognizable by their softness and ductility. An incrustation of OXIDE OF LEAD, PbO, *deep yellow when hot, light yellow when cold,* is formed upon the charcoal, not far from the place where the substance is heated.

Pure lead, or alloys containing a large amount of lead, when heated on charcoal, without soda, burn with a *blue* flame, and give the PbO incrustation.

Mix together acetate of lead, nitrate of silver, and mercurous nitrate, add dilute chlorhydric acid until a precipitate ceases to form on further addition, and filter the liquid. The precipitate contains all the *silver* and *mercury* and a part of the *lead* as chlorides. Add to the filtrate an equal bulk of alcohol and a small quantity of dilute sulphuric acid. All the lead which it contains will be precipitated. Filter, wash the precipitate, and test it for lead on charcoal.

Make a hole in the bottom of the filter and wash the precipitate obtained with chlorhydric acid into a small flask. Test the precipitate for the remainder of the lead, and for silver and mercury, according to Part III. (***67***), (***68***), and (***69***).

SECTION II.

MERCURIC COMPOUNDS, COPPER, BISMUTH, AND CADMIUM.

Metals which are not precipitated by chlorhydric acid.

MERCURIC COMPOUNDS.

$HgCl_2$.

Sulphydric Acid and Ammonic Sulphide precipitate mercuric compounds, at first as *double salts*, which appear first *white*, then *yellow*, *orange*, and *brown*, and finally as MERCURIC SULPHIDE, HgS, *black.* Mercuric sulphide does not dissolve when it is boiled with moderately concentrated nitric or chlorhydric acid. It dissolves readily in aqua regia.

Sodic Hydrate precipitates mercuric compounds at first as *basic salts*, *reddish brown*, finally as MERCURIC OXIDE, HgO, *yellow ;* insoluble in an excess of the precipitant.

Ammonic Hydrate precipitates mercuric compounds as SALTS CONTAINING AMMONIA, *white.* The precipitate is insoluble in an excess of ammonic hydrate.

Stannous Chloride reduces mercuric compounds, and precipitates them as MERCUROUS CHLORIDE, Hg_2Cl_2, *white.* After the metals of Group V., Section I., if they are present have been removed from a solution by the addition of chlorhydric acid, mercuric compounds are the only ones which give a precipitate in acid solution with stannous chloride.

Reactions with the Aid of Heat. Dry mercurous and mercuric chlorides form white sublimates when they are heated in a closed glass tube. All dry compounds of mercury, when they are mixed with dry sodic carbonate, and heated in a closed tube, give a sublimate of metallic mercury. The sublimate is at first a faint metallic film, which augments until drops of mercury appear. If the quantity of mercury is small, the film may be made to take the form of metallic drops by rubbing it with a copper wire.

COPPER.

$CuSO_4$; $CuCl_2$.

Metallic copper dissolves readily in dilute nitric acid. It dissolves with difficulty in chlorhydric acid. All copper solutions are *blue or green.*

Iron or Zinc precipitates copper from its acid solutions, either as a metallic coating or as *brownish red* metallic grains. If a strip of zinc and one of platinum are placed in a dilute acid solution of copper, so that they touch each other, the platinum is plated with copper.

Sulphydric Acid and Ammonic Sulphide precipitate cupric compounds as CUPRIC SULPHIDE, CuS, *black.* Cupric sulphide is insoluble in dilute acids, but dissolves readily in strong acids. It is somewhat soluble in an excess of ammonic sulphide.

Sodic Hydrate precipitates cupric compounds as CUPRIC HYDRATE, $Cu(HO)_2$, *light blue.* On boiling, CUPRIC OXIDE, CuO, *black*, is formed. The precipitate is insoluble in an excess of sodic hydrate.

Ammonic Hydrate precipitates cupric compounds as CUPRIC HYDRATE, which dissolves immediately in an excess of ammonic hydrate. The solution has a very intense *blue* color.

A valuable test for cupric compounds.

Ferrocyanide of Potassium precipitates cupric compounds in acid solutions, as FERROCYANIDE OF COPPER, $Cu_2(FeCy_6)$, *reddish brown.* This is a delicate test for very small quantities of copper.

Blowpipe Reactions. Cupric compounds color the borax bead *blue* when cold, and *green* when hot, in the oxidizing flame. The bead is colored *red*, and becomes *opaque* in the reducing flame. No other metal produces this coloration.

When cupric compounds are heated with sodic carbonate on charcoal, metallic copper is reduced in the form of small globules, which can be easily recognized by their *hardness and red color.*

Copper Flame. Compounds containing copper (alloys and salts) color the flame of a lamp *green*, or if chlorine is present, *blue.* By moistening a cupric compound with chlorhydric acid the *blue color* can easily be detected.

BISMUTH.

$Bi(NO_3)_3$.

Metallic bismuth dissolves readily in moderately concentrated nitric acid. It dissolves with great difficulty in chlorhydric acid.

Water. Solutions of bismuth, particularly those containing chlorhydric acid, are remarkable for giving a precipitate consisting of a BASIC SALT, when water is added to them, unless they contain a large quantity of free acid. The basic salt can be dissolved by the addition of an acid. CHLORHYDRIC ACID precipitates nitrate of bismuth solution as a BASIC CHLORIDE because the basic chloride of bismuth is more insoluble than the other basic salts. The precipitate is soluble on the further addition of chlorhydric acid.

Sulphydric Acid and Ammonic Sulphide precipitate bismuth compounds as the SULPHIDE OF BISMUTH, Bi_2S_3, *black.*

Sodic and Ammonic Hydrates precipitate bismuth compounds as the HYDRATE OF BISMUTH, $Bi(HO)_3$, *white.*

The hydrate of bismuth is insoluble in an excess of sodic and ammonic hydrates.

Blowpipe Reactions. Bismuth compounds, mixed with sodic carbonate, and heated on charcoal, give *brittle* metallic globules, and an incrustation of OXIDE OF BISMUTH, Bi_2O_3, on the charcoal.

The incrustation is *orange yellow* when hot, and *bright yellow* when cold.

CADMIUM.

Cd.

Metallic cadmium dissolves readily in nitric acid, and the solution contains nitrate of cadmium, $Cd(NO_3)_2$.

Sulphydric Acid and Ammonic Sulphide precipitate cadmium compounds as the sulphide, CdS, *lemon yellow.* This is insoluble in ammonic sulphide and potassium cyanide, but readily soluble in hot dilute sulphuric or nitric acid.

Sodic and Ammonic Hydrates precipitate compounds of cadmium as the hydrate, $Cd(HO)_2$, very readily soluble in a slight excess of ammonic hydrate, but insoluble in sodic hydrate.

When compounds of cadmium are mixed with sodic carbonate and exposed on charcoal to the inner blowpipe flame, the charcoal becomes covered with an incrustation of yellow or brownish-yellow oxide of cadmium, CdO.

Mix together mercuric chloride, cupric sulphate, and the nitrates of bismuth and cadmium, and add sulphydric acid until the metals are completely precipitated. Filter, and wash the precipitates with a little water; spread the filter out; scrape the precipitate from it, and heat the precipitate gently in a porcelain dish with strong nitric acid until red fumes cease to be given off; then add a little water and boil for a few minutes. The sulphide of *mercury* remains insoluble. Filter it off and test according to Part III. (***89***). Test the filtrate for *bismuth*, *copper*, and *cadmium* according to Part III. (***91***), (***92***), and (***93***).

GROUP VI.

TIN, ANTIMONY, ARSENIC, AND GOLD.

THE metals of Group VI. sometimes act as acids, uniting with metals, and sometimes as metals, uniting with acids. Their combinations with acids have an acid reaction; those with metals which are soluble in water have an alkaline reaction.

Sulphydric Acid precipitates the metals of Group VI. as SULPHIDES. The precipitation takes place slowly, particularly in the case of arsenic acid, and it should never be considered complete unless a current of sulphydric acid gas is passed through the solution for some time, and it is left to stand twenty-four hours.

The sulphides of metals of Group VI. are insoluble in dilute acids, but they are decomposed or dissolved by boiling with concentrated acids. They are soluble in sodic hydrate and in ammonic sulphide (the sulphide of gold with difficulty).

TIN.

Sn.

Metallic tin dissolves readily in strong chlorhydric acid on boiling, and the solution contains STANNOUS CHLORIDE, $SnCl_2$. Tin is oxidized to STANNIC OXIDE, SnO_2, *white powder*, by strong nitric acid. Stannic oxide is insoluble in nitric acid, but dissolves readily in hot concentrated chlorhydric acid, and the solution contains STANNIC CHLORIDE, $SnCl_4$. Stannic chloride is also formed by the action of aqua regia on metallic tin.

Metallic Zinc precipitates METALLIC TIN from its solutions in acids as crystalline metallic particles.

Blowpipe Reactions. Compounds of tin are reduced, when they are heated with sodic carbonate and potassic cyanide on charcoal ; and METALLIC TIN may be discovered in flattened globules by rubbing the fused mass, taken from the charcoal, in a mortar with water, and washing several times by decantation. When the globules are large, they may be observed on the charcoal during the fusion.

Oxidation by Sodic Nitrate. When a sulphide of tin is oxidized at the lowest possible temperature by a mixture of sodic nitrate and carbonate, STANNIC OXIDE, SnO_2, which is insoluble in water, is formed. If the oxidation is carried on at too high a temperature, stannate of sodium is formed, which is soluble in water.

STANNOUS COMPOUNDS.

$SnCl_2$.

Sulphydric Acid and Ammonic Sulphide precipitate stannous compounds from acid solutions as STANNOUS SULPHIDE, SnS, *dark brown.* Stannous sulphide dissolves with difficulty in a colorless solution of ammonic sulphide (the mono-sulphide), and it is scarcely soluble in ammonic hydrate and carbonate. It is converted into stannic sulphide by a yellow solution of ammonic sulphide (the poly-sulphide), and dissolves readily when warmed with the solution. Stannous sulphide is soluble in sodic hydrate. The sulphides of tin are precipitated from these solutions, when a dilute acid is added gradually, until the reaction becomes strongly acid.

Sodic and Ammonic Hydrates and Carbonates precipitate stannous compounds as STANNOUS HYDRATE, $Sn(HO)_2$, *white.*

Stannous hydrate dissolves in an excess of sodic hydrate, but it is insoluble in an excess of sodic carbonate and of ammonic hydrate and carbonate.

Mercuric Chloride changes stannous chloride, $SnCl_2$, into STANNIC CHLORIDE, $SnCl_4$, and a *white* precipitate of MERCUROUS CHLORIDE, Hg_2Cl_2, is formed. *No other metal in solution gives this reaction with mercuric compounds.*

STANNIC COMPOUNDS.

$SnCl_4$.

Sulphydric Acid and Ammonic Sulphide precipitate stannic compounds as STANNIC SULPHIDE, SnS_2, *light yellow.* Stannic sulphide dissolves readily in ammonic sulphide, and in sodic hydrate, and is precipitated completely from the solution, when a dilute acid is added gradually until the reaction becomes strongly acid. Stannic sulphide is nearly insoluble in ammonic carbonate.

Sodic and Ammonic Hydrates and Carbonates precipitate stannic compounds as STANNIC HYDRATE, $Sn(HO)_4$, *white.* Easily soluble in sodic hydrate.

Mercuric Chloride does not give a precipitate with stannic chloride.

Metallic Zinc precipitates TIN as crystalline metallic particles from stannic compounds in an acid solution, and the tin can be easily recognized by dissolving the metallic particles, after they have been washed by decantation. They are dissolved by heating them with a few drops of strong chlorhydric acid, and the stannous chloride thus obtained gives the precipitate, above described, with mercuric chloride.

ANTIMONY.

Sb.

Metallic antimony is scarcely attacked by chlorhydric acid. It is oxidized when heated with moderately strong nitric acid to ANTIMONIC OXIDE, Sb_2O_5, which is almost completely insoluble in nitric acid, but is readily soluble in hot concentrated chlorhydric acid. Aqua regia dissolves metallic antimony as ANTIMONIC CHLORIDE, $SbCl_5$.

Water precipitates solutions of antimony containing chlorhydric acid, particularly those of antimonious compounds, as a BASIC CHLORIDE. The precipitation can be prevented, or the precipitate dissolved by the addition of a sufficient quantity of acid; for this purpose tartaric acid is the most suitable.

Metallic Zinc partially precipitates METALLIC ANTIMONY from its acid solutions, and when a piece of platinum in contact with a piece of zinc is introduced into solutions of antimony containing an excess of chlorhydric acid, metallic antimony is deposited upon the platinum as a *dark-brown stain. No other metal produces the same stain under like circumstances.*

Antimoniuretted Hydrogen. To obtain this body, follow exactly the directions given for obtaining arseniuretted hydrogen. (See page 59.)

The mirror obtained with antimoniuretted hydrogen consists of a *black*, *sooty*, metallic deposit. It dissolves very slowly in hypochlorite of sodium. If the deposit is moistened with *yellow* ammonic sulphide, an *orange-yellow* stain appears on the spot when it is dried. These reactions are unimportant as tests for antimony, but a knowledge of them is necessary, in order that they may not be mistaken for evidences of the presence of arsenic.

Blowpipe Reactions. Antimony compounds, mixed with sodic carbonate and potassic cyanide, are quickly reduced, at a comparatively low temperature, to METALLIC ANTIMONY, *brittle shining grains.* The metal is completely volatilized by a strong heat. Metallic antimony, when heated on charcoal, burns with a white smoke, and gives an incrustation which is deposited at some distance from the place heated, and is very volatile before the blowpipe flame.

Oxidation by Sodic Nitrate. When a sulphide of antimony is oxidized by a mixture of sodic carbonate and nitrate, ANTIMONIATE OF SODIUM, Na_3SbO_4, insoluble in water, is formed.

ANTIMONIOUS COMPOUNDS.

$SbCl_3$; $KSbOC_4H_4O_6$, *Tartar Emetic.*

Sulphydric Acid and Ammonic Sulphide precipitate antimonious compounds as ANTIMONIOUS SULPHIDE, Sb_2S_3, *orange red.* The precipitation should be made in a cold solution containing tartaric acid and very little free chlorhydric acid. Antimonious sulphide dissolves in sodic hydrate and in ammonic sulphide, most readily in *yellow* ammonic sulphide, and is precipitated completely from these solutions when a dilute acid is added gradually until the reaction becomes acid. It is nearly insoluble in ammonic carbonate. *Yellow* ammonic sulphide converts it into ANTIMONIC SULPHIDE, Sb_2S_5. (See below.)

Sodic and Ammonic Hydrates and Carbonates precipitate antimonious compounds as ANTIMONIOUS OXIDE, Sb_2O_3, *voluminous white precipitate.* The precipitation only takes place after the lapse of a considerable time in solutions containing tartaric acid. Antimonious oxide is soluble in sodic hydrate.

ANTIMONIC COMPOUNDS.

$SbCl_5$; Sb_2O_5.

Sulphydric Acid and Ammonic Sulphide precipitate antimonic compounds in acid solution as ANTIMONIC SULPHIDE, Sb_2S_5, *orange red.* The precipitation does not take place immediately ; but first an *orange* color appears ; and it is only after passing sulphydric acid for a long time through the solution that all the antimonic sulphide is precipitated. The precipitation should be made in a cold solution containing tartaric acid and very little free chlorhydric acid. Antimonic sulphide has the same properties as antimonious sulphide.

Antimonic oxide, Sb_2O_5, plays the part of an acid with bases, and forms insoluble compounds with SODIUM, and soluble compounds with POTASSIUM.

Compounds of antimony can best be recognized by the stain, which they form on platinum, when the metal is precipitated by zinc from their solutions. (See page 57, Antimony.)

ARSENIC.

As.

Metallic arsenic is readily oxidized to arsenious or to arsenic compounds by nitric acid and is dissolved. It is not dissolved by chlorhydric acid. Metallic arsenic volatilizes at a heat below redness in a tube or on charcoal, and produces an odor like garlic, very characteristic of compounds of arsenic. The same odor is produced by heating a dry compound of arsenic on charcoal.

Arseniuretted Hydrogen. To obtain this body, and to use Marsh's test, provide a 4 oz. flask with a tight-fitting cork, into which a funnel-tube and a small tube drawn off to a point and bent at a right angle * are introduced. Fill it two-thirds full of dilute sulphuric acid, and add several pieces of pure zinc. A brisk evolution of hydrogen only commences after a few minutes. When that point is reached, wait five minutes for the expulsion of the air contained in the flask (without this precaution there is danger of an explosion), and light the hydrogen issuing from the point of the bent tube. The opening of the point and the quantity of gas evolved should be such that the hydrogen burns with a blunt flame

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* In order to dry the gas, it is better to fit to the cork a chloride of calcium tube, bent downwards, so that it will not tip the flask over by its weight, and to adapt to the chloride of calcium tube, by means of a cork, a suitable jet directed upwards.

about $\frac{1}{4}$ in. in length. Hold a bit of porcelain in the flame, in order to be certain that the reagents employed are free from arsenic. (See below.) If this is the case, pour a little of an acid liquid containing arsenic into the funnel-tube, while the hydrogen flame continues lighted. The flame after a few moments becomes white, and leaves a black stain of metallic arsenic upon a cold porcelain object held in it, in the same way that the flame of a candle would leave a deposit of soot on a cold surface. The arsenic stain or mirror is *shining black* (not *sooty black*, like antimony). It dissolves quickly in hypochlorite of sodium. When the stain is moistened with yellow ammonic sulphide solution, and the spot is dried, it is *bright yellow.*

Reaction with the Aid of Heat. All dry compounds of arsenic, except some compounds with the higher metals, when they are mixed with dry sodic carbonate and potassic cyanide, and heated in a sealed tube, give a sublimate of metallic arsenic, which can be best recognized by breaking the sealed end of the tube after the formation of the sublimate, and by heating the sublimate quickly until it begins to volatilize, and by smelling of the upper end of the tube. A smell of garlic is proof of the presence of arsenic.

ARSENIOUS COMPOUNDS.

As_2O_3.

Arsenious oxide is sparingly soluble in water It dissolves more readily in sodic hydrate or carbonate, or in chlorhydric acid.

Sulphydric Acid and Ammonic Sulphide precipitate arsenious compounds from acid solutions as ARSENIOUS SULPHIDE, As_2S_3, *yellow.* Arsenious sulphide dissolves very readily in sodic hydrate and ammonic sulphide, and it also dissolves, although less readily, in ammonic carbonate.

It is precipitated from these solutions, when a dilute acid is slowly added until the reaction becomes acid.

Sodic and Ammonic Hydrates and Carbonates produce no precipitate in arsenious compounds.

ARSENIC COMPOUNDS.

As_2O_5.

By heating any compound of arsenic with strong nitric acid, or by fusing a dry compound with a mixture of dry sodic carbonate and nitrate in a porcelain crucible, arsenic acid, H_3AsO_4, or arseniate of soda, Na_3AsO_4, is formed. These compounds are soluble in water.

Sulphydric Acid and Ammonic Sulphide do not immediately precipitate acid solutions of arsenic acid.

The solution (which should contain free chlorhydric acid) at first becomes yellow, when a current of sulphydric acid is passed through it, and finally a *yellow* precipitate is formed; reduction to the state of an arsenious compound takes place slowly, and at the end of several days all the arsenic is precipitated as ARSENIOUS SULPHIDE, As_2S_3, *yellow*. This operation may be very materially accelerated by boiling the solution.

See the properties of arsenious sulphide above.

Argentic Nitrate produces no precipitate in acid solutions of arsenic compounds. (When chlorhydric acid is present a white precipitate of argentic chloride forms, and must usually, after the addition of an excess of argentic nitrate, be separated by decantation and filtration.) If sodic hydrate is added to a clear solution containing arsenic acid, and containing also an excess of argentic nitrate, until a permanent precipitate forms, and if then acetic acid is added until the reaction becomes acid, ARGENTIC ARSENIATE, Ag_3AsO_4, *brick red*, is precipitated. This precipitate is soluble in ammonic hydrate and in dilute nitric acid.

This is the usual test for arsenic.

GOLD.

$AuCl_3$.

Metallic gold is insoluble in any single acid, but it dissolves readily in aqua regia.

Ferrous Sulphate solution in considerable quantity precipitates METALLIC GOLD, Au, *purplish-brown powder*, from acid solutions containing gold.

Stannous Chloride solution precipitates METALLIC GOLD, Au, *purple flakes, very finely divided*, from acid solutions containing gold. (A very dilute solution of stannous chloride should be added, drop by drop, to the gold solution.)

Sulphydric Acid gas slowly precipitates gold compounds in acid solutions as SULPHIDE OF GOLD, Au_2S_3, *brown flakes*. The precipitate is soluble after a long digestion with *yellow* ammonic sulphide.

Blowpipe Reaction. Gold is easily separated from its compounds with non-metallic elements by heating them on charcoal. It can be recognized as bright yellow globules.

Gold can easily be separated from the other metals and recognized by its precipitation with ferrous sulphate, or by its insolubility in any single acid.

Mix together solutions containing tin, antimony, and arsenic, dilute the solution and add a considerable quantity of dilute chlorhydric acid. Pass sulphydric acid through the solution for an hour, and set it over night in a warm place. Filter the precipitate, wash and dry it completely. It contains all the metals in the form of sulphides; separate them according to Part III. (*83*) and the following tests.

TESTS FOR ACIDS.

GROUP I.

ARSENIOUS, ARSENIC, CHROMIC, SULPHURIC, PHOSPHORIC, BORACIC, OXALIC, FLUORHYDRIC, CARBONIC, AND SILICIC ACIDS.

Acids which are precipitated from neutral or slightly alkaline solutions by baric chloride.

SECTION I.

ARSENIOUS, ARSENIC, AND CHROMIC ACIDS.

Acids which are precipitated as sulphides or reduced to an oxide by sulphydric acid.

ARSENIOUS and ARSENIC acids are precipitated as sulphides, and must always be detected by sulphydric acid.

Arsenic with this reagent plays the part of a metal. (See pages 60 and 61.)

CHROMIC ACID.

H_2CrO_4 ; $K_2Cr_2O_7$.

Solutions containing chromic acid have a *yellow* color. Those which have an acid reaction are of a deeper color, and very concentrated acid solutions are *red*. Solutions containing chromic acid can usually be recognized by these colors. Chromic acid can easily be reduced to chromic oxide (see page 38) by boiling its solution with chlorhydric acid and alcohol, or by adding to the solution sulphydric acid or ammonic sulphide. Therefore, in all solutions to which these latter reagents have been added, chromic acid is changed to chromic oxide, and must be looked for among the metals.

Baric Chloride precipitates chromic acid in neutral solutions, or in solutions of which the only free acid is acetic acid (*i. e.*, solutions which have been made alkaline by sodic hydrate, and acid by acetic acid), as BARIC CHROMATE, $BaCrO_4$, *light yellow*. Chromium may be discovered in the precipitate of baric chromate thus obtained, by dissolving some of it in the borax bead. The *green color* of chromic oxide appears, even though very little chromate of barium was contained in the precipitate.

No other acid which precipitates baric chloride gives the same color in the borax bead.

SECTION II.

SULPHURIC AND SULPHUROUS ACIDS.

Acids which are precipitated by baric chloride in acid solutions, either immediately or after oxidation.

Blowpipe Reactions. All compounds containing sulphur, when they are pulverized, moistened, and mixed with sodic carbonate, and when the mixture is heated on charcoal, form sodic sulphide. The sodic carbonate must be heated until it soaks into the charcoal, then if the portion of the charcoal which has absorbed it is dug out with a knife, moistened with water, and laid on a piece of bright silver, the presence of sulphur may be detected by the appearance of a black stain of sulphide of silver. If no silver is at hand, the sodic sulphide can be extracted by soaking with water the portion of the charcoal which has been heated. The solution, after it has been filtered, and acidified with acetic acid, gives a precipitate of SULPHIDE OF LEAD, PbS, *black*, with a solution of ACETATE OF LEAD.

For the separation of sulphuric and sulphurous acids, see Part III. **(123)** *and* **(124)**.

SULPHURIC ACID.

H_2SO_4; Na_2SO_4.

Baric Chloride precipitates compounds of sulphuric acid in neutral or acid solutions as BARIC SULPHATE, $BaSO_4$, *white powder*. Baric sulphate is insoluble in chlorhydric, nitric, sulphuric, and the weaker acids. No other acid gives a like precipitate with $BaCl_2$, in acid solutions. COMPOUNDS OF LEAD AND STRONTIUM and STANNIC OXIDE *are the only others precipitated by sulphuric acid.*

Calcic Chloride in concentrated solutions gives a precipitate of CALCIC SULPHATE, $CaSO_4$, but calcic sulphate is soluble in 380 parts of water.

SULPHUROUS ACID.

H_2SO_3; Na_2SO_3.

Sulphurous acid, in acid solutions, can be recognized by the smell of sulphurous oxide which is given off. Sulphurous oxide is evolved from acid solutions of sulphites when they are heated.

Baric Chloride precipitates compounds of sulphurous acid in neutral solutions as BARIC SULPHITE, $BaSO_3$, *white powder*. Baric sulphite is decomposed and dissolved by chlorhydric and nitric acids. POTASSIC DICHROMATE oxidizes sulphurous acid in a solution, made acid with chlorhydric acid, to SULPHURIC ACID, $2(H_2CrO_4) + 3(H_2SO_3) = Cr_2(SO_4)_3 + 5H_2O$.

Sulphydric Acid decomposes sulphurous acid in acid solutions with precipitation of sulphur.

$$2H_2S + H_2SO_3 = S_3 + 3H_2O.$$

The Iron Reduction Test. POTASSIC FERRICYANIDE AND FERRIC CHLORIDE, in consequence of the reduction of the ferric chloride, give a *blue* precipitate when these solutions are mixed, and a drop of the mixture is held on the end of a glass rod in an atmosphere containing sulphurous acid. Sulphydric acid gives the same reaction, and when sulphydric acid is present it is necessary to add to the solution sufficient plumbic acetate to precipitate it, before applying this test for sulphurous acid. This reaction is produced equally well in a solution ; but there are many other reducing agents which act in the same manner, but which do not take the gaseous form, and for this reason it is best to add chlorhydric acid to the solution, to warm it, and to test the gas evolved for sulphurous acid with potassic ferricyanide and ferric chloride.

Sulphurous acid is usually recognized by its smell, or by the above test.

SECTION III.

PHOSPHORIC, BORACIC, OXALIC, AND FLUORHYDRIC ACIDS.

Acids which are precipitated by baric chloride in neutral solutions, but which are not precipitated in acid solutions.

For the tests by which these acids can be most easily detected, see Part III. (**125-128**).

PHOSPHORIC ACID.

H_3PO_4 ; Na_2HPO_4.

All the Metals of Groups II., III., IV., V., precipitate compounds of phosphoric acid in neutral or in slightly alkaline solutions as PHOSPHATES. The precipitate is soluble in acids. The phosphoric acid cannot usually be separated from the metal by treating the precipitate with ammonic hydrate, and from most metals it can only be partially separated by a treatment with sodic hydrate or carbonate. MAGNESIC SULPHATE solution, to which a solution of ammonic chloride and of ammonic hydrate is added, precipitates compounds of phosphoric acid as MAGNESIC AMMONIC PHOSPHATE, $MgNH_4PO_4$, *white crystalline powder*, soluble in acids but insoluble in ammonic hydrate. When the quantity of phosphoric acid is very small, the precipitate only forms after some time. *This is the best reagent for phosphoric acid in neutral or slightly alkaline solutions.*

Molybdate of Ammonium* precipitates compounds of phosphoric acid in an acid solution as a *yellow crystalline phospho-molybdate of ammonium.* No other acid gives a like precipitate. *The best reagent for phosphoric acid in acid solutions is molybdate of ammonium, when there are bases present which are precipitated by ammonic hydrate.*

Sulphydric acid and ferrocyanhydric acid also precipitate ammonic molybdate solution. (See Part III., ***127.***)

BORACIC ACID.

H_3BO_3; $Na_2B_4O_7$, *Borax.*

All the Metals of Groups II., III., IV., V. precipitate compounds of boracic acid in neutral or slightly alkaline solutions as BORATES. The precipitate is soluble in acids.

The precipitation of boracic compounds by metals of these groups is mostly incomplete, and the acid can be separated from the metal in almost all cases by treating the precipitate with ammonia or with sodic hydrate or carbonate.

Flame Test. Boracic acid in dry compounds or in con-

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* It is best to pour a few drops of the solution to be tested into the ammonic molybdate solution, instead of adding the ammonic molybdate to the solution to be tested.

centrated solutions imparts a beautiful *green color* to the flame of burning alcohol.

In order to perform the test, mix the dry boracic acid compound with a few drops of strong sulphuric acid in a small evaporating dish, add alcohol, and set the alcohol on fire. Stir the contents of the dish constantly during the combustion of the alcohol, and observe the *green color*, when the alcohol has mostly burned away.

Turmeric Paper Test.*[*]* Boracic acid compounds, when their solution has been rendered acid by chlorhydric acid, turn a piece of turmeric paper, which has been dipped in the solution and completely dried, *brownish red. This is the usual test for boracic acid.*

OXALIC ACID.

$H_2C_2O_4$; KHC_2O_4.

All the Metals of Groups II., III., IV., V. precipitate salts of oxalic acid in neutral solutions as OXALATES, except chromic oxide compounds. The precipitate is soluble in acids, and in the case of many metals the oxalic acid is removed from the precipitate by ammonic hydrate and by sodic hydrate and carbonate.

Sulphate of Calcium precipitates oxalic acid and its salts in a solution to which sufficient ammonic hydrate to render the solution strongly alkaline, and then sufficient acetic acid to render the reaction acid, have been added, as CALCIC OXALATE, CaC_2O_4, *white powder. No other acid, except fluorhydric acid, produces a precipitate under these circumstances, and calcic fluoride cannot easily be mistaken for calcic oxalate.* (*See fluorhydric acid, below.*)

Sulphuric Acid (concentrated) decomposes dry compounds, or highly concentrated solutions of compounds of

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[*] If iron is present it is necessary to boil the solution with sodic carbonate in excess, to filter, and to use the filtrate for the turmeric paper test.

oxalic acid, with evolution of CARBONOUS and CARBONIC OXIDES. *Effervescence takes place.* $H_2C_2O_4 = CO + CO_2 + H_2O$.

FLUORHYDRIC ACID.

HF; CaF_2; NH_4F.

Fluorhydric acid cannot be present in acid solutions in glass vessels.

BARIUM, STRONTIUM, and CALCIUM salts in neutral solutions precipitate fluorhydric acid as BARIUM, STRONTIUM, or CALCIUM Fluoride, BaF_2, SrF_2, CaF_2. The two former are voluminous white precipitates. Calcium fluoride is a gelatinous transparent precipitate, whose formation it is very difficult to observe.

It is usually unnecessary to test for fluorhydric acid except in solid substances.

Sulphuric Acid (concentrated) sets fluorhydric acid free from its solid compounds, and the acid may be recognized by its property of etching glass. To perform the test, mix the pulverized substance, containing fluorine, with strong sulphuric acid in a lead cup,* or in a platinum crucible. Prepare a piece of glass by melting wax on it and pouring off all that does not adhere, leaving a thin coating of wax on the surface; scratch lines in the wax, laying the surface of the glass bare; cover the vessel containing the fluorine compound and sulphuric acid with the glass, and warm gently. After fifteen minutes warm the glass, and rub off the wax; the surface exposed by the scratches will be etched by the fluorhydric acid.

Fluorhydric acid must always be removed, if present, by heating with strong sulphuric acid, before the other acids of Group I., except sulphuric acid and carbonic acid, are tested for in a substance.

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* A lead cup may be made by hammering up the end of a piece of one-inch lead pipe until it is entirely closed, and by sawing off the pipe 1½ inches from the end. Such a cup may be warmed on a sand-bath sufficiently for making the test without danger of melting it.

SECTION IV.

CARBONIC ACID AND SILICIC ACID.

Acids which are precipitated by baric chloride in neutral solution, but which are set free by acids, and which cannot be present in a solution that has been evaporated with an excess of an acid.

CARBONIC ACID.

Na_2CO_3 ; K_2CO_3.

Carbonic acid can only be present in considerable quantity in a solution which has an *alkaline reaction.*

All the Metals of Groups II., III., IV., V. precipitate alkaline carbonates.

Every Acid sets free CARBONIC DIOXIDE, CO_2, from a solution of a carbonate. An effervescence or the formation of bubbles can be observed when an acid is added to a solution of a salt of carbonic acid. Carbonates insoluble in water are also decomposed by a free acid, carbonic dioxide being evolved.

Carbonic dioxide can easily be recognized by the white precipitate of CARBONATE OF CALCIUM, $CaCO_3$, which the gas produces in a drop of lime-water,* held in it on the end of a glass rod.

SILICIC ACID.

K_2SiO_3.

Silicic acid combined with bases in a solution is set free by all acids, and being decomposed into water and silicic oxide ($H_2SiO_3 = H_2O + SiO_2$), the latter often separates as a precipitate ; frequently, however, it remains for months in a solution after an acid has been added.

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* The lime-water is soon destroyed by the absorption of carbonic dioxide from the air, and before using it to test for carbonic acid, the student should assure himself that a drop of it gives a precipitate with the gas evolved from sodic carbonate, to which chlorhydric has been added.

When silicic oxide, or silica, has been separated from its combination with a base by the addition of an acid, and the solution has been evaporated completely to dryness, the silica remains perfectly insoluble when the dry mass is treated with water or acids to dissolve the bases. *This is the characteristic test for silicic acid.*

When silicic acid is present in a solution it is recognized and separated by the above process before any other tests are performed.

Many solid compounds of silicic acid are not acted upon by acids, and can only be brought into solution by the process described, Part III., XXIII.

The following properties of silicic acid must be considered, in order to determine whether a body, which has been left insoluble after a treatment with an acid, is silica, or some other compound, which is likewise insoluble:

Silicic acid is precipitated from its solutions by acids in the *gelatinous form*, or in the form of *amorphous white flakes*. Silica which has been dried always has the latter form.

Silica, after it has been fused with four parts of a mixture of sodic carbonate and potassic carbonate, forms a glass, which is entirely soluble in water. When the solution thus obtained is evaporated with an excess of nitric acid, and the dry mass is treated with water, no metal should go into solution which gives a precipitate with sodic carbonate.

GROUP II.

CHLORHYDRIC, BROMHYDRIC, IODHYDRIC, CYANHYDRIC, FERROCYANHYDRIC, FERRICYANHYDRIC, AND SULPHYDRIC ACIDS.

Acids which are not precipitated by baric chloride, but which are precipitated by argentic nitrate in nitric acid solution.

SECTION I.

CHLORHYDRIC, BROMHYDRIC, IODHYDRIC, AND CYANHYDRIC ACIDS.

Acids which give with argentic nitrate a white or light yellow flocculent precipitate, insoluble in dilute nitric acid, and are not precipitated by salts of iron in acid solution.

CHLORHYDRIC ACID.

HCl ; $NaCl$.

Plumbic, Mercurous, and Argentic Salts are the only compounds which give, with chlorhydric acid, precipitates insoluble in nitric acid.

Sulphuric Acid (concentrated) sets free chlorhydric acid from its compounds. Chlorhydric acid gas precipitates a drop of argentic nitrate, held on the end of a glass rod in an atmosphere containing it, as ARGENTIC CHLORIDE, $AgCl$, *white flakes.* Compounds containing cyanhydric, chloric, and hypochlorous acids, produce the same reaction. Chlorhydric acid gas does not bleach indigo solution. See, however, Part III. (**48**).

Argentic Nitrate precipitates solutions containing chlorhydric acid as ARGENTIC CHLORIDE, AgCl, *white flakes*, turning *purple* on exposure to light, *settling quickly after they have been shaken.* Argentic chloride is soluble in ammonic hydrate, and completely insoluble in boiling nitric acid (concentrated). No other compound of silver except the ferro- and ferricyanide remains undissolved after this treatment. When ferro- or ferricyanhydric acid is present in a solution containing chlorhydric acid, follow the directions given, Part III. (**136**) (*a*), before applying the argentic nitrate test.

This is the usual test for chlorhydric acid.

BROMHYDRIC ACID.

KBr.

Nearly all bromides are soluble in water ; bromide of lead, however, dissolves very sparingly, and the mercurous and argentic salts are quite insoluble.

Argentic Nitrate precipitates from solutions of bromides the ARGENTIC BROMIDE, AgBr, *white*, closely resembling the chloride, but more difficultly soluble in ammonia ; insoluble in hot nitric acid.

Bromides are decomposed by *chlorine*, *hypochlorites*, strong *sulphuric* and *nitric acids*, bromine being set free, which imparts a yellow or yellowish-red color to the liquid. On shaking the tube with a little ether or carbonic disulphide the bromine will be dissolved and impart to the solvent a yellow or reddish-brown color.

IODHYDRIC ACID.

KI.

The iodides resemble very much the corresponding chlorides and bromides.

Argentic Nitrate produces in solutions of iodides a *yellowish-white* precipitate of ARGENTIC IODIDE, AgI, very slightly soluble in ammonia, but soluble in boiling concentrated nitric acid.

Mercurous Nitrate precipitates *yellowish-green* MERCUROUS IODIDE, Hg_2I_2.

Mercuric Chloride precipitates brilliant *scarlet* MERCURIC IODIDE, HgI_2.

Chlorine or bromine water liberates iodine from iodides ; on shaking with carbonic disulphide the iodine is concentrated in this liquid, forming a violet-colored solution.

Free iodine imparts a blue color to starch paste. For this purpose dilute starch paste is added to a solution of an iodide, and then *chlorine water*, *nitric acid*, or *sulphuric acid* carefully added to liberate the iodine. If too much chlorine is added chloride of iodine is formed, which prevents the formation of the blue color. The chloride of iodine is also produced by nitric acid in presence of considerable amount of chlorides.

CYANHYDRIC ACID.

HCy ; KCy.

The reactions of different classes of cyanhydric acid compounds must be considered separately.

Soluble Simple Cyanides. Cyanides of metals of Groups I., II., and III. are soluble in water. Cyanhydric acid is set free from their solutions by even the feeblest acids (acetic and carbonic). (MERCURIC CYANIDE is soluble in water, but is not decomposed by alkalies nor by acids, except by sulphydric acid ; and the tests described for cyanhydric acid cannot be applied to it. Sulphydric acid precipitates mercuric sulphide, and sets cyanhydric acid free.)

Insoluble Simple Cyanides. Cyanides of metals of Groups IV. and V., except mercuric cyanide, are insoluble in water, and the cyanides of metals of Group V. are not decomposed, or are decomposed with great difficulty by acids. The insoluble cyanides dissolve readily in potassic cyanide, and the ordinary tests for metals cannot be used with such solutions. The cyanides can be precipitated from these solutions by the addition of an acid, with some exceptions, the two most remarkable of which are described separately. (See Ferro- and Ferricyanhydric Acids.)

Free Cyanhydric Acid can be recognized by its smell, which is like that of bitter almonds. (The acid is very poisonous, and the fumes arising from a solution containing a considerable quantity of it should be inhaled with caution.)

Argentic Nitrate precipitates soluble compounds of cyanhydric acid in an acid solution as ARGENTIC CYANIDE, AgCy, *white flakes, which do not settle so readily, when shaken, as argentic chloride, and which do not turn purple quickly in the light.* Argentic cyanide is wholly decomposed and dissolved by boiling a few minutes with strong nitric acid. It is also decomposed, and cyanhydric acid goes into solution, when it is digested with dilute chlorhydric acid in contact with metallic zinc.

Prussian Blue Test. When a feebly acid solution containing cyanhydric acid is mixed with several drops of ferrous sulphate solution, and with a drop of ferric chloride solution, and sodic hydrate is added until a precipitate forms, and the mixture is warmed for a minute, and then acidified with dilute chlorhydric acid, a *blue precipitate*, or more frequently a *blue coloration*, appears, either immediately or after the addition of a drop of ferric chloride.

When ferro- or ferricyanhydric acid or both acids are present (see the following section), they must be removed from the solution before the Prussian-blue test can be applied. To this end add to a small quantity of the solution an equal bulk of dilute sulphuric acid, and dilute with a considerable quantity of water ; add ferric chloride or ferrous sulphate, or both together, according as ferro- or ferricyanhydric acid or both acids are present, and then add baric chloride until the blue precipitate appears of a much lighter shade; shake thoroughly, and allow the precipitate to settle for a few minutes, and filter. If only the first few drops run through the filter blue, they should be thrown away and the remainder of the filtrate taken. If no clear filtrate can be obtained, add to the filtrate a little baric chloride and filter again. The filtrate is to be tested as above by the addition of sodic hydrate, and afterwards of an acid for cyanhydric acid. A sufficiently capacious flask must be chosen for the operation. The only object in adding baric chloride is to facilitate the filtration from the blue precipitate.

Cyanhydric acid is the only acid which gives this reaction under these circumstances.

SECTION II.

FERROCYANHYDRIC AND FERRICYANHYDRIC ACIDS.

Acids which give with argentic nitrate colored precipitates, which are not wholly destroyed on boiling with strong nitric acid, and which are precipitated by ferrous or ferric salts, and by cupric salts in dilute acid solutions.

FERROCYANHYDRIC ACID.

$H_4(FeCy_6)$; $K_4(FeCy_6)$.

Ferrous Sulphate precipitates ferrocyanhydric acid compounds in acid solutions as POTASSIC FERROUS FERROCYANIDE, $K_2Fe(FeCy_6)$, *bluish white* precipitate, which quickly turns *dark blue* through oxidation by the air.

Ferric Chloride precipitates ferrocyanhydric acid compounds in acid solution as PRUSSIAN BLUE, $Fe_4(FeCy_6)_3$, *deep blue.*

Ferrocyanhydric acid is the only acid which gives this reaction.

Cupric Sulphate precipitates ferrocyanhydric acid compounds in acid solution as CUPRIC FERROCYANIDE (Cu_2FeCy_6), *brownish-red powder.*

The metals are left as oxides, and the ferrocyanogen is dissolved as sodic ferrocyanide, when these precipitates are digested with sodic hydrate.

FERRICYANHYDRIC ACID.

$H_6(Fe_2Cy_{12})$; $K_6(Fe_2Cy_{12})$.

Ferrous Sulphate precipitates compounds of ferricyanhydric acid in acid solution as TURNBULL'S BLUE, $Fe_3(Fe_2Cy_{12})$, *deep blue.*

Ferricyanhydric acid is the only acid which gives this reaction.

Ferric Chloride does *not* precipitate compounds of ferricyanhydric acid in acid solution. The color of the solution is deepened.

Cupric Sulphate precipitates compounds of ferricyanhydric acid in acid solution as CUPRIC FERRICYANIDE, *yellowish-green powder.*

The metals are left as oxides, and the cyanogen is dissolved as sodic ferricyanide when these precipitates are treated with sodic hydrate.

SECTION III.

SULPHYDRIC ACID.

An acid which gives a black precipitate with salts of lead, silver, copper, and many others in an acid solution.

No other acid gives a precipitate of the same color with these metals.

SULPHYDRIC ACID.

H_2S ; $(NH_4)_2S$.

Sulphydric Acid is set free from its solutions by all other acids except carbonic and cyanhydric acid, and it can be recognized by its smell. The sulphydric acid gas is given off with effervescence when the solution is concentrated.

The Metals of Groups I. and II. form with sulphydric acid soluble sulphides, which have an alkaline reaction.

The Metals of Group IV., when the acid with which they are combined is neutralized, form with sulphydric acid INSOLUBLE SULPHIDES, which, with the exception of the sulphides of cobalt and nickel, are dissolved by cold dilute chlorhydric acid, with evolution of sulphydric acid.

Metals of Groups V. and VI. form with sulphydric acid insoluble sulphides, which are not decomposed by dilute acids. (See also Mercury, page 48 and page 50.)

Lead-Paper Test. A piece of paper moistened with plumbic acetate, and held over a solution from which sulphydric acid is set free by the addition of a stronger acid, is *blackened. No other acid gives this reaction.*

GROUP III.

NITRIC, CHLORIC, AND ACETIC ACIDS.

Acids which are not precipitated by any metal.

SECTION I.

NITRIC AND CHLORIC ACIDS.

Acids which deflagrate when tested with the blowpipe on charcoal.

NITRIC ACID.

HNO_3 ; $NaNO_3$.

Nitric Acid, when concentrated, is readily decomposed when heated with copper turnings, and *red* fumes of NITRIC PEROXIDE, NO_2, are given off. The reaction can be obtained with a moderately dilute solution by adding to it concentrated sulphuric acid. No reaction is obtained with very dilute solutions.

Ferrous Sulphate Test. Add a few drops of a solution containing nitric acid to concentrated sulphuric acid in a test-tube, and pour upon this solution a layer of cold ferrous sulphate solution. A *brown* or *red* color appears at the line of separation of the two solutions, arising from the absorption of nitrous gases by the ferrous sulphate.

This is the characteristic test for nitric acid.

CHLORIC ACID.

$KClO_3$.

Sulphuric Acid (concentrated). When a small quantity of a solid chlorate, or a very concentrated solution containing a chloric acid compound, is added to strong sulphuric acid, and heat is applied, *a peculiar yellow gas* (oxides of chlorine) is evolved, which has a characteristic suffocating odor, which precipitates ARGENTIC CHLORIDE in a drop of an argentic nitrate solution, and which bleaches a drop of an indigo solution when these reagents are held on the end of a glass rod in an atmosphere containing the gas.

This is the characteristic test for chloric acid.

Hypochlorous Acid gives the same reactions as chloric acid, but that acid is easily set free and evolved from its solution by dilute sulphuric acid, while chloric acid is not, and moreover it is usually present only in alkaline solutions.

Chlorhydric Acid in the presence of an oxidizing agent gives a similar reaction, but the yellow gas evolved (chlorine) is much less intense in color, and has a different odor. It is, however, very difficult to distinguish between the reaction given by chlorine in such a case and that given by chloric acid compounds.

SECTION II.

ACETIC ACID.

An acid which does not deflagrate on charcoal.

ACETIC ACID.

$HC_2H_3O_2$; $NaC_2H_3O_2$.

The Strong Mineral Acids set acetic acid free from its combinations.

Acetic acid can be recognized by the *odor of vinegar* peculiar to it.

Sulphuric Acid Test. When an equal bulk of alcohol is added to strong sulphuric acid, and a small quantity of a solution containing a compound of acetic acid is added, and the mixture is heated, a characteristic odor of acetic ether is given off.

In case gases are given off, which make it difficult to recognize the odor of acetic ether, it is advisable to provide the test-tube in which the reaction is performed with a tube for distillation,* and to distil a small quantity of the alcohol into another test-tube, to mix the distillate with water, to neutralize it with sodic carbonate, and to warm it; the odor of acetic ether can then be recognized in the liquid which was distilled.

This is the characteristic test for acetic acid.

Argentic and Mercurous Nitrates precipitate concentrated neutral solutions of acetic acid compounds as ARGENTIC AND MERCUROUS ACETATES, $AgC_2H_3O_2$ and $Hg_2(C_2H_3O_2)_2$, *white crystalline scales.* The precipitates are soluble in dilute nitric acid, and also in a large quantity of water.

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* Bend a 3-16 inch tube, of about one foot in length, at an angle of about 80°, so that one arm shall only be 1½ inches long. Fit a cork to the test-tube, and insert the bent tube in a hole bored through the cork with a round file.

PART III.

PRELIMINARY TESTS WITH NON-METALLIC SOLIDS.

EXAMINATION IN A CLOSED TUBE.

Use a piece of hard glass tubing three-eighths of an inch in diameter, closed at one end (see page 26) for this examination. Introduce the substance, pulverized or in small pieces, into the tube, wipe the inside of the tube if necessary with a bit of rolled filter-paper, and heat the substance, gently at first, but eventually to the highest temperature attainable with the flame of a Bunsen's lamp or with the blowpipe flame. Observe carefully the changes which occur.

(1) ***No Change.*** The substance contains no organic matter, no readily fusible body, no readily volatile body, and no water.

Pass to the Examination on Charcoal (page 85).

(2) ***Water.*** Substances containing water (usually water of crystallization) deposit a *film of moisture* in the upper part of the tube when they are heated. If the water colors turmeric paper *brown*, AMMONIA is present.

(3) ***Organic Matter.*** Substances containing organic matter *blacken* and give off gases when they are heated.

Should the substance contain organic matter it must be burnt, until the organic matter is completely destroyed,* by heating with the lamp or blowpipe, on

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* In some special cases, as in examinations for mercury and arsenic, other processes of analysis must be employed, for which larger works must be consulted.

platinum foil or on a bit of porcelain, or in a porcelain dish or crucible, before further analysis, commencing with the examination on charcoal (page 85), is proceeded with.

A GAS IS GIVEN OFF.

Oxygen. May be recognized by its property of rekindling
(*4*) a glimmering match held in the tube.

PEROXIDES, NITRATES, and CHLORATES evolve oxygen.

Nitrates and chlorates also deflagrate on charcoal. See **(*14*)**.

Sulphurous Oxide, SO_2, can be recognized by its smell.
(*5*) Some SULPHATES of higher metals, and many SULPHITES, evolve sulphurous oxide when they are heated.*

Sulphydric Acid, H_2S, can be recognized by its smell
(*6*) and by its property of blackening lead paper. See **(*45*)**.

Some alkaline SULPHIDES, containing water, evolve sulphydric acid when they are heated.

Carbonic Dioxide, CO_2, can be recognized by its prop-
(*7*) erty of extinguishing a lighted or glimmering match held in the tube. See also **(*41*)**.

Some CARBONATES lose carbonic dioxide when they are heated.

Hyponitric Oxide, NO_2, appears as *red* fumes.
(*8*) NITRATES of the higher metals evolve hyponitric oxide when they are heated.

Ammonia, NH_3, can be recognized by its smell, and by its

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* Many sulphides of higher metals give off sulphur in the form of sulphurous oxide when they are roasted with access of air. The sulphides, finely pulverized, may be heated red-hot in a tube, open at both ends, and held in an inclined position to favor the draught, and the sulphurous oxide may be detected by its smell at the upper end of the tube.

(9) property of turning moist turmeric paper *brown.* Salts of ammonia, in the presence of alkalies, and some organic substances, evolve ammonia when they are heated.

[Better tests for these bodies, with the exception of oxygen, are given in the following pages, since it is often difficult to observe the formation of a gas in a small tube ; the phenomena described above should, however, be looked for when substances are heated in a closed tube.]

A SUBLIMATE FORMS.

An opinion may be formed of the volatility of the sublimate, according to the distance from the heated part of the tube at which it is deposited.

Sulphur sublimes easily and solidifies in *reddish-brown* **(10)** drops, which become *yellow* or *yellowish-brown* on cooling.

Some METALLIC SULPHIDES give off a portion of their sulphur when they are heated.

Ammonic Salts form white sublimates. Touch the sublimate **(11)** with a drop of sodic hydrate, or with a bit of paper moistened with sodic hydrate, and if the smell of ammonia is given off it consists of an ammonic salt.

Mercury. Metallic mercury sublimes as a *grey film*, which **(12)** augments to form *globules* when the quantity of mercury is large.

MERCURIC SULPHIDE, HgS, gives a *black* sublimate, which becomes *red* when it is rubbed.

MERCUROUS CHLORIDE, Hg_2Cl_2, and MERCURIC CHLORIDE, $HgCl_2$, give a *white* sublimate, which turns *black* when it is moistened with ammonic sulphide solution.

Arsenic. Metallic arsenic sublimes and deposits itself as a **(13)** *brilliant black metallic ring* in the tube.

Arsenious oxide, As_2O_3, forms a *white crystalline* sublimate, which turns *yellow* when it is moistened with sulphydric acid solution.

Arsenious sulphide, As_2S_3, forms a sublimate, which is *reddish yellow* when hot, and *yellow* when cold. It is somewhat less volatile than sulphur.

RECAPITULATION (*10*) TO (*13*) SUBLIMATES.

The substance is heated in a closed tube.

White sublimate—ammonic salts (***11***); mercurous chloride, Hg_2Cl_2, and mercuric chloride, $HgCl_2$ (***12***); and arsenious oxide, As_2O_3 (***13***).

Yellow sublimate—sulphur (***10***); and arsenious sulphide, As_2S_3 (***13***).

Brown sublimate (while hot)—sulphur (***10***).

Reddish-yellow sublimate (while hot)—arsenious sulphide, As_2S_3 (***13***).

Gray metallic sublimate—mercury (***12***).

Black sublimate—arsenic (***13***); and mercuric sulphide, HgS, red when rubbed (***12***).

EXAMINATION ON CHARCOAL.

Hollow out a small cavity in a piece of charcoal (see page 25), and heat a portion of the solid substance with the blowpipe flame.

Nitrates and Chlorates enter into a vivid combustion, called *deflagration*, when they are heated on charcoal.
(***14***)

Potassium and Sodium Salts melt, and some of (***15***) them are imbibed by the pores of the charcoal when they are heated.

Compounds of the Metals of Groups II. and (***16***) ***III., also Zinc Compounds and Silicic Oxide,*** remain as a *white infusible* mass on the charcoal after heating. Frequently, when heat is first applied, they melt in their water of crystallization, and afterwards become solid.

ALUMINIC OXIDE becomes *blue*, and ZINC OXIDE becomes *green* when they are moistened with cobaltic nitrate and heated in the oxidizing flame.

Salts of the Metals of Groups IV. and V. leave a (***17***) dark-colored residue when they are heated on charcoal. The oxides of these metals generally assume a darker color when they are heated. Exceptions: Zinc and mercury.

Salts of Ammonia and Mercury, also Com- (***18***) ***pounds of Arsenic and Antimony,*** which do not contain another metal, volatilize completely when they are heated on charcoal.

Gold and Silver Compounds, also Oxides of (***19***) ***Lead and Bismuth,*** give *bright metallic* globules when they are heated on charcoal.

FUSION WITH SODIC CARBONATE.

(***20***) When metals of Groups IV., V., and VI. appear to be present (see ***17***), mix a small quantity of the pulverized substance with two or three times its bulk of sodic carbonate in the palm of the hand, moisten with water, and form the mixture by working it with a knife-blade into a ball the size of a pea. Place the ball in a cavity scooped out of a piece of charcoal, and heat with the inner blowpipe flame until almost all of the carbonate of soda has been imbibed by the charcoal. Many metals are reduced and appear as metallic globules in the cavity of the charcoal, and those which are volatile deposit an incrustation of their oxides on the charcoal. This incrustation is to be looked for at a greater or less distance from the cavity, according to the volatility of the metal, and always in the direction in which the metallic vapors are blown by the flame.

The physical and chemical properties of the globules and the color of the incrustations afford means of recognizing several metals, usually, however, only when they are not associated with others.

Iron, Cobalt, Nickel, and Manganese Compounds give neither globule nor incrustation. (21)

Gold, Silver, and Copper Compounds give *malleable globules*, which can be distinguished by the respective colors of the metals. They give no incrustation. (22)

Zinc Compounds give no globules, but a *white incrustation*, ZnO, near the spot heated. The incrustation is *yellow* while hot. It is not volatile in the oxidizing flame. It becomes *green* when it is moistened with nitrate of cobalt, and heated in the oxidizing flame. (23)

Tin Compounds give very *ductile white globules.*

(24) The incrustation, SnO_2, produced by tin compounds is *dirty yellow* when hot, and *lighter* when cold. It is deposited in the immediate vicinity of the cavity, and it is very difficult to distinguish it from the ash of the charcoal.

Lead Compounds give very *ductile globules.*

(25) The incrustation, PbO, is *bright yellow* when hot, and *pale yellow* when cold. It is deposited at a greater distance than SnO_2 from the cavity. When the blow-

pipe flame is directed upon the incrustation of PbO it vanishes, and the flame is colored *blue*.

Bismuth Compounds give *brittle globules*. The incrus-
(**26**) tation, Bi_2O_3, is *orange yellow* when hot, and *bright yellow* when cold. It vanishes when the blowpipe flame is directed upon it, but it does not impart a *blue* color to the flame.

Cadmium Compounds give no globules, but a *yellow* to
(**27**) *reddish-brown* incrustation, very different in color from that of any other metal.

Antimony Compounds give *brittle globules*, but metallic
(**28**) antimony is so volatile that frequently these are driven off by the heat required for their reduction. Sometimes fumes, arising from the vapor of antimony, are visible. The incrustation, Sb_2O_3, is *white*. It is deposited at a greater distance than PbO from the cavity, and it can easily be driven from one place to another on the charcoal by the heat of the blowpipe flame.

Arsenic Compounds give no globules, but a character-
(**29**) istic *garlic odor*. The incrustation, As_2O_3, is *white*, and it is still more volatile than Sb_2O_3.

When the compound contains several metals that can be reduced, they alloy with each other, and it is usually impossible to recognize the metals in the presence of each other by their physical properties; also, the incrustation given by one metal frequently obscures that given by another.

If a sufficient quantity of the metal can be easily reduced, it is always advisable to treat it with solvents in the manner to be described under metals. (See page 98.)

Sulphur. The following modification of the fusion with
(**30**) carbonate of sodium on charcoal is a valuable test to discover sulphur in baric sulphate and in sulphides. If a piece of the charcoal which has imbibed the soda is moistened and laid on a silver coin, a *black* stain appears, if the coin is washed after a few minutes,

in case sulphur is present. The charcoal may also be pulverized and treated with water, and if the solution, after being filtered, gives a black precipitate with plumbic acetate solution, sulphur is present. For this test the flame of a candle, oil lamp, or alcohol lamp must be used, else the sulphur in the burning gas will vitiate the results.

RECAPITULATION OF THE EXAMINATION ON CHARCOAL.

The substance is heated on charcoal.

DEFLAGRATION.—Nitrates and chlorates (***14***).

FUSION.—Potassium and sodium salts (***15***).

WHITE INFUSIBLE RESIDUE.—Compounds of metals of Groups II. and III., zinc salts and silicic acid (***16***).

DARK-COLORED RESIDUE.—Compounds of metals of Groups IV. and V., except zinc and mercury (***17***).

COMPLETE VOLATILIZATION.—Ammonic and mercuric compounds, and compounds of arsenic and antimony, which contain no other metal (***18***).

BRIGHT METALLIC GLOBULES.—Silver and gold compounds, and the oxides of lead and bismuth (***19***).

The substance is mixed with sodic carbonate and heated on charcoal.

METALLIC GLOBULES WITHOUT INCRUSTATION.—Gold, silver, and copper (***22***); tin (***24***).

METALLIC GLOBULES AND INCRUSTATION.—Lead (***25***); bismuth (***26***); antimony (***28***).

INCRUSTATION WITHOUT GLOBULES.—Zinc (**23**); Cadmium (**27**). GARLIC ODOR.—Arsenic (**29**).

FORMATION OF SODIC SULPHIDE.—All compounds containing sulphur (**30**).

PRELIMINARY TESTS WITH METALLIC BODIES.

EXAMINATION IN A CLOSED TUBE.

See page 26.

Mercury. (**31**) Amalgams containing mercury give a sublimate of METALLIC MERCURY when they are heated. At first a *gray film* forms in the upper part of the tube, and, when the amount of mercury is considerable, *fine globules* of metallic mercury are formed, which agglomerate and become more distinctly visible when they are rubbed with a copper wire.

Arsenic. (**32**) Some metallic compounds, containing ARSENIC, give a metallic *mirror* or ring in the upper part of the tube when they are heated.

EXAMINATION ON CHARCOAL.

(**33**) Heat a piece, one-fourth as large as a pea, of the metallic substance in a cavity on a piece of charcoal. See Part I., page 25. The phenomena to be observed are the formation of the incrustations described (pages 87 and 88), the smell of ARSENIC, and the vapors of MERCURY and ANTIMONY.

COPPER colors the blowpipe flame *green*, or in the presence of chlorine *blue*.

PRELIMINARY TESTS WITH NON-METALLIC SOLIDS (*continued*).

TESTS WITH THE BORAX BEAD.

If the substance to be tested appears to be the oxide, or an oxygen-salt of a higher metal (see ***17***), dissolve some of it in the borax bead.

Cobalt colors the bead *blue* in the oxidizing and in the re-
(***34***) ducing flame.

Copper colors the bead *green* when hot, and *blue* when cold,
(***35***) in the oxidizing flame. It colors the bead *red*, when cold, in the reducing flame.

Chromium colors the bead *green* in both flames.
(***36***)

Iron colors the bead *brownish red* when hot, and *yellow* when
(***37***) cold, in the oxidizing flame.

Nickel colors the bead *violet* when hot, and *pale brown* when
(***38***) cold, in the oxidizing flame. The color disappears in a *good* reducing flame. (See page 44.)

Manganese colors the bead *amethyst* in the oxidizing flame.
(***39***) The color disappears in a good reducing flame.

(***40***) THE OXYGEN COMBINATIONS OF THE REMAINING METALS color the bead very slightly or not at all.

CONCENTRATED SULPHURIC ACID TEST.

CONCENTRATED SULPHURIC ACID, with the aid of heat, sets free other acids from most of their combinations with metals, and frequently in such a form that they can be recognized by simple tests. This reaction is not, of course, a method of separation; one acid may obscure the test for another, and the possible cases are so complicated that it would be useless to attempt to describe them all; therefore, if the result of the sulphuric acid test appears doubtful, it is best to reserve judgment of its value until after the tests for acids in solution have been applied.

The reactions of acids or their salts, when a small quantity of a solid substance or of a very concentrated solution is added to a few cubic centimetres of strong sulphuric acid in a test-tube are described below. Heat should be applied after the reaction, which takes place at the ordinary temperature, has been observed.

Carbonic Acid. *Effervescence.* Carbonic dioxide gas **(41)** renders turbid a drop of lime-water held on the end of a glass rod in the test-tube (see page 70, foot-note). Carbonic acid is also detected by the chlorhydric acid test, and in many cases that test is preferable to the one with sulphuric acid, since oxalic acid does not give the same reaction with chlorhydric acid. See **(72)**.

Oxalic Acid. When a dry compound of oxalic acid is **(42)** added to strong sulphuric acid, and the mixture is heated, carbonous oxide and carbonic oxide are evolved.

The sulphate of calcium test for oxalic acid **(125)** is more accurate than that with sulphuric acid.

Cyanhydric, Ferro- and Ferricyanhydric **(43)** **Acids.** Compounds of these acids evolve, when perfectly dry, carbonic oxide with effervescence when they are heated with strong sulphuric acid. Usually, however, a faint odor of cyanhydric acid can be detected. The special tests are more valuable. See **(73)**, and **(133)**, **(134)**, and **(135)**.

Fluorhydric Acid. When a solid substance or a con- **(44)** centrated solution containing fluorine is heated with strong sulphuric acid, fluorhydric acid is evolved, which etches glass (see page 69); when silicic acid or a silicate is present, fumes of fluoride of silicon, which give a precipitate of silica in a drop of water held over them on the end of a glass rod, are evolved.

If fluorhydric acid is discovered, the substance which is to be used for page 104, III., and the following tests, must be heated with sulphuric acid in a platinum vessel until the fluorhydric acid is driven off completely.

A separate portion can be used for testing for sulphuric acid. See (***30***).

Sulphydric Acid. Compounds containing this acid (***45***) evolve it (often with effervescence) when strong sulphuric acid is added to them. Sulphydric acid may be recognized by its smell and by its property of blackening paper dipped in plumbic acetate solution. See (***74***) and (***129***).

Sulphurous Acid. Compounds containing this acid (***46***) evolve sulphurous oxide, SO_2, with effervescence when strong sulphuric acid is added to them. Sulphurous oxide, when free from sulphydric acid, and from some others, can be recognized by its smell. See (***75***).

Chloric and Hypochlorous Acids. (See page 80.) (***47***) Compounds containing these acids evolve a *yellow* gas on the addition of strong sulphuric acid, even when the mixture is not heated. The gas can best be recognized by its color, its odor, and its strong bleaching action on a drop of indigo solution held in the tube on the end of a glass rod. This gas precipitates nitrate of silver.

Chlorhydric Acid. Compounds containing chlorhydric (***48***) acid evolve the gas, HCl, frequently with effervescence, on the addition of sulphuric acid. The gas does not bleach indigo solution, but precipitates nitrate of silver held on the end of a glass rod in the tube. Chlorhydric acid, in the presence of an oxidizing agent, evolves chlorine under the same circumstances. The gas produces the same reactions as are produced by the gas evolved by chloric and hypochlorous acids, but it can be distinguished from them by its smell and by its color, which is a less intense yellow. For special test, see (***130***).

Bromhydric Acid. Bromides are decomposed by strong (***49***) sulphuric acid with evolution of bromhydric acid, which, if the sulphuric acid is concentrated and in excess, is partly decomposed, with separation of bromine and formation of sulphurous oxide. (See also page 73.)

Iodhydric Acid. All the iodides are decomposed by (***50***) strong sulphuric acid on the application of heat. Iodine is set free, which escapes in violet vapors and imparts a blue color to paper moistened with starch. (See also page 73.)

Nitric Acid. Compounds containing nitric acid in con- (***51***) siderable quantity produce reddish fumes when heated with sulphuric acid in the presence of copper turnings or of any other reducing agent. The following test for nitric acid is more delicate: Mix a little of the powdered substance or solution with strong sulphuric acid, and pour cautiously upon the acid a solution of ferrous sulphate. A *brown* or *red* color at the line of separation of the two solutions indicates the presence of NITRIC ACID, HNO_3. See (***136***) (*a*) for this test in the presence of ferro- or ferricyanhydric acid.

Acetic Acid gives with sulphuric acid an odor of vine- (***52***) gar. The following test is more delicate: Add an equal volume of alcohol to strong sulphuric acid, and then add the solid substance or concentrated solution supposed to contain ACETIC ACID, $C_2H_4O_2$. If this acid is present, the odor of acetic ether will be perceptible on heating the mixture. It is well to add pure acetic acid at the same time to a mixture of sulphuric acid and alcohol, in order to compare the odor produced with that observed in the test. If other gases render the odor of acetic ether difficult to perceive, the precautions described on page 81 must be observed.

RECAPITULATION OF SULPHURIC ACID TEST.

A colorless gas is given off.

THE GAS IS WITHOUT ODOR.—Carbonic acid (***41***); oxalic acid (***42***).

A PUNGENT SUFFOCATING ODOR.—Fluorhydric acid (***44***); chlorhydric acid (***48***); bromhydric acid (***49***).

AN ODOR OF BITTER ALMONDS.—Cyanhydric, ferrocyanhydric, and ferricyanhydric acids (***43***).

AN ODOR OF ROTTEN EGGS.—Sulphydric acid (***45***).

AN ODOR OF BURNING SULPHUR.—Sulphurous acid (***46***).

AN ODOR OF VINEGAR.—Acetic acid (***52***).

A colored gas is given off.

The gas has also a peculiar suffocating odor.

YELLOW GAS.—Chloric and hypochlorous acids (***47***).

FAINT YELLOW GAS.—Chlorine (***48***).

RED FUMES.—Nitric acid with copper (***51***).

VIOLET FUMES.—Iodhydric acid (***50***).

METHODS OF DISSOLVING NON-METALLIC BODIES.

Substances to be tested may be divided into three classes.

1st Class: Bodies soluble in water.

(***53***) To ascertain whether a body is entirely soluble, take a few grains of the substance, which, when it dissolves with difficulty, must be in as finely divided a condition as possible, and digest them with a considerable quantity of water in a test-tube. If the substance does not dissolve completely, boil it for a few minutes with water. If it still does not dissolve completely, filter and evaporate a few drops of the filtrate on platinum foil, in order to see whether anything has dissolved.

If the substance is partly soluble in water, treat a considerable quantity in the manner directed above, repeat the boiling with water once or twice, wash thoroughly with water, and filter, taking care that as little as possible of the solid substance goes upon the filter, and use the insoluble portion for the following test.

A separate analysis should usually be made of the part of a substance which is soluble, and of that which is insoluble in water.

If the substance is wholly insoluble in water, it may be used immediately for the following test. It should usually be finely pulverized.

2d Class: ***(54)*** Bodies insoluble in water, but soluble in chlorhydric acid or in nitric acid or in aqua regia. Digest a few grains, or a very small quantity of the finely powdered substance with dilute chlorhydric acid. If it does not entirely dissolve, boil it with the acid. If it still does not entirely dissolve, pour off the liquid, and boil the insoluble substance with strong chlorhydric acid. If the substance is not wholly soluble in chlorhydric acid, repeat the trial with nitric acid in the same way, using a fresh portion of the substance. If the substance is still insoluble, use a mixture of both acids (aqua regia).

If sulphur, which may be recognized by its color and its low specific gravity, or silicic acid, which may be recognized by its peculiar gelatinous aspect, separate out on the addition of chlorhydric or of nitric acids, the sub-

stance must be considered as soluble, and after filtration (in case of silicic acid, see **64**) *the solution must be tested according to page* 104 (*III.*), *or page* 107 (*VI.*), *according as nitric or chlorhydric acid has been the solvent.*

If the substance is only partially soluble after treatment as above, the insoluble portion must be washed carefully and separated by filtration from the soluble portion, and it must be subjected to the tests for the 3d Class. See page 137.

In all cases before proceeding to test the solution obtained by treatment with acids, it must be so prepared that it will be moderately acid with chlorhydric acid, and will not contain any free nitric acid, or great excess of strong acid of any kind. It the solution was made by the use of dilute chlorhydric acid it is already in this condition. If nitric acid was used for the solution the excess must be replaced by chlorhydric acid. In this case add a quantity of the latter about half as great as that of the nitric acid used ; evaporate the liquid almost to dryness in the hood, and add from 25 to 50 c.c. of water to the residue. If the substance was dissolved in concentrated chlorhydric acid, the solution should be evaporated down in the same manner to expel the excess of that acid, and the residue diluted considerably with water. If the addition of chlorhydric acid should cause a precipitate in the clear acid solution, or if the residue, after evaporation, does not dissolve entirely in water, it can only be occasioned by the presence of argentic, plumbic, or mercurous chloride. In this case, after the addition of water, filter off the insoluble powder and examine for those metals as directed (**68**), (**67**), (**69**). Use the filtrate for the sulphydric acid test (**76**).

3d Class: Bodies insoluble in water, and in chlorhydric, (**55**) nitric, and nitro-chlorhydric acids (aqua regia).

The tests to be applied to bodies of this class follow those in the scheme for testing bodies in solution. See page 137.

METHODS OF DISSOLVING METALLIC BODIES.

Metals are divided into three classes.

1st Class: Metals which are not attacked by NITRIC ACID.

(*56*) If the metal does not appear to be entirely soluble in dilute nitric acid, even after boiling, use strong nitric acid.

GOLD is insoluble, also alloys containing a very large proportion of gold.

2d Class: Metals which are attacked by nitric acid, and

(*57*) converted into oxides (*white powder*), which are insoluble in the acid.

TIN and ANTIMONY belong to Class II.

3d Class: Metals which dissolve entirely in nitric acid.

(*58*) All the remaining metals dissolve entirely in nitric acid. With metals of this class the solution should be effected by heating with nitric acid in an evaporating dish until no more red fumes are given off. Chlorhydric acid should then be added, the greater part of the acid evaporated, water added, and the solution should be tested according to (page 104, III.). If the metal remains partly insoluble, a small portion of it should be carefully tested by boiling with strong nitric acid, to see whether it will not dissolve by using a sufficient quantity of the acid.

(*59*) An alloy may contain metals of each class, therefore, after the treatment with nitric acid described in the last paragraph, if an insoluble residue is found it should be dissolved in aqua regia, and the solution boiled until the smell of chlorine ceases to be given off. Then if the residue had a metallic appearance the solution must be tested for gold. See (***87***). If it was a white powder it must be tested for tin and antimony. See (***86***) and (***85***).

SOLUTION IN CHLORHYDRIC ACID.

(***60***) Sometimes a test shows that the metallic body can be readily dissolved in chlorhydric acid (see ZINC, page 39, and IRON, page 41), and in this case the chlorhydric acid solution should be preferred, and it should be tested according to page 107, VI.

TESTS FOR METALS AND ACIDS.

BODIES IN SOLUTION.

CLASSES I. AND II. (See page 95.)

Substances dissolved in water or in acids.

(The tests I. and II., which follow, are only to be used for substances dissolved in water, or where the solvent is unknown.)

I.—REACTION WITH TEST-PAPER.

(61) Observe the reaction with litmus paper.

(62) If the reaction is acid, add to a small portion of the solution sodic carbonate, drop by drop, until the effervescence ceases, and then heat to boiling.

If metals of the 2d, 3d, 4th, and 5th groups are present, a precipitate will be formed, except in a few special cases.

Some idea of the amount of free acid present in the solution can be formed by observing the amount of sodic carbonate required to neutralize it.

II.—EVAPORATION.

(63) Test one or two drops of the solution by evaporation on platinum foil (or on a bit of glass or porcelain, if there is danger of injury to the platinum), to see whether it leaves a residue.

If this test shows the presence of solid matter in the solution, a portion of the latter may be evaporated to dryness in a small porcelain dish, and some of the tests for solids, particularly the EXAMINATION ON CHARCOAL (page 85), and the CONCENTRATED SULPHURIC ACID TEST (page 91), may be applied to the dry substance thus obtained. It is best only to use the evaporation and tests applied to the dry substance, to settle any doubts that remain in regard to the constitution of the solution, after the tests usually applied to solutions have been performed. (See page 104, etc.)

When a solution of unknown origin is presented for analysis it should always be heated, in order to see whether a gas is given off that can be recognized by the tests (page 92, and page 93).

Silicic Acid **(*64*)** is usually not present, except in alkaline solutions; it occurs, however, in small quantities in spring and river waters, and it may also exist in acid solutions.

Unless silicic acid is known to be absent it should be tested for, and removed before making the remaining tests. To this end render the solution acid with chlorhydric acid, and evaporate carefully to dryness in a small porcelain dish. Care must be taken not to heat the dish over the lamp after its contents have become dry. Treat the dry substance with a few drops of acid, and then boil with water. If there is an insoluble residue, it consists of silica, SiO_2. (See page 71.)

Subject the chlorhydric acid solution to the remaining tests, beginning with page 107, VI.; and if there is reason to suspect that the precipitate, left after the treatment with an acid, contains other substances besides silica, examine it according to page 105, IV.

TESTS FOR METALS.

THE following scheme of testing for metals is founded upon the *successive* precipitation of a number of groups, which include all the metals as far as Group II. After the metals of the higher groups have been removed by precipitation, or have been found to be absent, those of Group II. are precipitated successively. Sodium and potassium are detected by the colors of their flames in a solution which has been freed from all the higher metals except magnesium. Ammonia can be detected in a solution without having reference to its other constituents.

After the separation into groups has taken place by precipitation with the general reagents, each precipitate, which may contain one or all the metals belonging to its group, is usually dissolved, and the further analysis is performed by testing in the several solutions for all the metals which they may contain. These special tests sometimes require the separation of the different metals, one after another, in a particular order, while sometimes a test for a metal may be applied to the solution without regard to the presence of other metals. In all cases the conditions requisite for applying the tests will be described.

The general tests must be applied in the following order: 1st, Chlorhydric acid to effect the precipitation of silver and mercurous compounds, and of lead if it is present in large quantity. 2d, Sulphydric acid in an acid solution to precipitate small quantities of lead and the metals of Group V., Section II., and of Group VI. (From this precipitate the metals of Group VI. are separated by dissolving them in ammonic sulphide.) 3d, Ammonic hydrate until the reaction becomes alkaline, ammonic chloride and ammonic sulphide to precipitate the metals of Groups III. and IV.

Whenever a single metal or a group of metals is precipitated some of the liquid containing the precipitate must be poured on a filter, and the first drops of the solution which run through must be tested with some of the reagent which was used to produce the precipitation, in order to ascertain whether it has been completely effected. Should a fresh precipitate make its appearance, everything must be poured back from the filter into the test-tube or flask and more of the reagent must be added. This operation must be repeated until no precipitation is produced by the same reagent in the liquid which runs through the filter. After a little practice it is easy to estimate how much of a reagent is required to effect a complete precipitation. A surplus over this quantity is called an excess of the reagent.

When it has been found that a slight excess of the reagent has been added, the whole of the liquid and the precipitate together must be poured upon a fresh filter and the liquid must be allowed to drain off. The liquid is to be tested for metals of the succeeding groups, and the precipitate must usually be completely freed from it; otherwise the separation has no value. To this end water must be blown on the precipitate from the wash-bottle and allowed to drain off. Care must be taken not to let the water overflow the edge of the filter. The washing must be continued until the water which flows through the filter is proved, either by evaporation on platinum foil or by the application of tests for the succeeding groups, not to contain any metal in solution. The last portions of the wash-water which pass through the filter may be thrown away, as they contain very little of the substances to be tested for.

The value of analyses depends upon the care with which the separation of precipitates from the liquid in which they have been formed is executed. For special directions for filtering, see pages 23 and 24.

When the reaction of a liquid in a test-tube is to be tested, always close the tube and shake it thoroughly before dipping the test-paper in.

METALS OF GROUPS VI. AND V.

III.—CHLORHYDRIC ACID TEST.

Metals in acid solution.

GROUP V., SECTION I.

In case the solution is known to contain chlorhydric acid, pass to page 107 (VI.).

If the solution has an alkaline reaction, pass to page 105 (IV.).

(65) Add to a very small quantity of the solution to be tested a few drops of dilute chlorhydric acid. If a precipitate forms, continue to add chlorhydric acid, drop by drop, as long as it seems to increase in quantity, then add a quantity of chlorhydric acid about equal to that already added.

If no precipitate is formed, or if it is dissolved on further addition of dilute chlorhydric acid, LEAD IN LARGE QUANTITY, SILVER, AND MERCUROUS SALTS ARE ABSENT. *Pass to* page 107 (VI.).

(66) If a permanent precipitate is formed, it may consist of PLUMBIC, ARGENTIC, AND MERCUROUS CHLORIDES. Most of the succeeding tests for the detection of bases must be performed in a solution freed from these metals. Therefore if a precipitate has been observed, treat a considerable quantity of the solution in the same way that the small portion was treated, and after an excess of chlorhydric acid has been added shake the liquid for a minute or two. The precipitate will then settle in a short time, leaving the solution nearly clear. The solution should be decanted through a filter, and the precipitate washed twice by decantation through the filter with water acidulated with chlorhydric acid.

Test the filtrate according to page 107 (VI.).

Lead.* (*67*)** Add a small quantity of water to the precipitate, and boil, then let the precipitate settle, and decant the clear liquid through the same filter which was used in (66***) into another vessel. Add to the filtrate an equal bulk of alcohol and a small quantity of dilute sulphuric acid. If a white precipitate forms it consists of SULPHATE OF LEAD, $PbSO_4$.

If lead is found, the precipitate must be washed as before, with boiling water by decantation, until the filtrate gives no *black* precipitate with ammonic sulphide. *If no precipitate remains after the washing*, NO ARGENTIC OR MERCUROUS SALTS *are present. Pass to* page 107 (VI.).

***Silver.* (*68*)** If a precipitate remains after washing with boiling water, add to it ammonia, pour the solution through a filter, and acidify with nitric acid. If a precipitate forms, it consists of ARGENTIC CHLORIDE, AgCl.

***Mercurous Salts.* (*69*)** If a precipitate remains after ammonia has been added, it will have a *gray or black color.* The precipitate consists of A MERCUROUS COMPOUND OF AMMONIA.

IV.—CHLORHYDRIC ACID TEST.

Metals in alkaline solutions.

(*70*) Add chlorhydric acid until the reaction becomes distinctly acid, and if a precipitate forms, wash it thoroughly with *cold* water upon a filter until the filtrate is no longer acid to test-paper. Observe whether sulphydric acid is given off.

If a precipitate is formed with chlorhydric acid, filter and test the filtrate according to page 107 (VI.).

(***71***) *If no precipitate is formed, pass to* page 107 (VI.).

(*a*) If the precipitate is *white*, it may consist of:

Plumbic Chloride. Boil a little of it with water, and test one portion of the solution for LEAD according to (***67***), and test another portion for CHLORINE by adding nitric acid and argentic nitrate.

Plumbic Sulphate. Test according to (***140***).

Argentic Chloride. Test according to (***141***).

(*b*) If the precipitate is *colored* it may contain

The Sulphides of Arsenic, Antimony, and Tin. Test according to page 109 (VIII.).

Sulphur may be precipitated, accompanied by a disengagement of sulphydric acid. The precipitated sulphur can be recognized by its appearance, and its insolubility in aqua regia.

V.—CHLORHYDRIC ACID TEST FOR ACIDS.

See also Silicic Acid (***64***).

If a solution is acid, many of the following tests, particularly (***73***), (***74***), and (***75***), can be applied by simply heating the solution.

Carbonic Acid is only present in alkaline solutions. CO_2
(***72***) is evolved with effervescence, when an acid is added, until the solution has an acid reaction. Hold a drop of lime-water on the end of a glass rod in the tube; if CARBONIC DIOXIDE, CO_2, is present, *a white precipitate* forms. (See page 70, foot-note.) The same test can be applied to solid carbonates. See also the sulphuric acid test (***41***).

The following acids need only be looked for when an *odor can be perceived after heating the solution, or after adding chlorhydric acid and heating.*

Cyanhydric Acid, in its soluble combinations with (***73***) most metals, is set free by chlorhydric acid. It can be recognized by its smell. See also the prussian blue test (***133***).

Sulphydric Acid is evolved from alkaline solutions (***74***) (often with effervescence), on the addition of chlorhydric acid, when the solution is heated. It can be recognized by its smell and by the lead paper test. (See the sulphuric acid test (***45***) and the argentic nitrate test (***129***).

Sulphurous Oxide is evolved from alkaline, neutral, or (***75***) slightly acid solutions of sulphites, on the addition of chlorhydric acid. Mix a little potassic ferricyanide and ferric chloride, and hold a drop of the mixture on the end of a glass rod in the tube after chlorhydric acid has been added and the tube heated. If a *blue* color appears, SULPHUROUS ACID is present. If sulphydric acid is present add sufficient plumbic acetate to precipitate it before performing the test. See the sulphuric acid test (***46***).

VI.—SULPHYDRIC ACID TEST.

Metals in acid solutions.

(***76***) Add sulphydric acid solution to a small quantity of the solution to be tested, and warm gently. In case metals of Group VI. are to be tested for, it is better to pass sulphydric acid gas into the dilute solution, made acid with chlorhydric acid. The total precipitation of the metals of Group VI. is frequently only effected after one or two days.

If no precipitate forms, no metals of Groups V. and VI. are present. Pass to page 115 (X.).

(*77*) If a precipitate forms observe the color. It may consist of the SULPHIDES OF LEAD, PbS; BISMUTH, Bi_2S_3; COPPER, CuS; MERCURY, HgS; and GOLD, Au_2S_3, when it is *black;* ARSENIC, As_2S_3; TIN (BISULPHIDE), SnS_2; CADMIUM, CdS, *yellow;* TIN (MONOSULPHIDE), SnS, *brown;* ANTIMONY, Sb_2S_3 or Sb_2S_5, *orange.* The presence of a *black sulphide* hides the color of the other sulphides, so that all may be present when the precipitate is black.

If only a light, fine, white precipitate, which is not destroyed by acids, is formed, it consists of sulphur, and is frequently due to the presence of a ferric salt or a chromate in the solution. In case only sulphur is precipitated, pass to page 115 (X.).

(*78*) If a precipitate forms in a small portion of the solution, a sufficient quantity for use in all the succeeding tests for metals must be treated with sulphydric acid until the metals of Groups V. and VI. are completely precipitated as sulphides; and the precipitate thus obtained must be washed on a filter quickly, with warm water containing sulphydric acid, until the addition of ammonic hydrate to the filtrate ceases to produce a precipitate, and it must then be treated according to (*79*).

Test the filtrate for metals of Groups IV., III., II., and I. (See page 115, X., etc.)

VII.—SOLUBILITY OF THE SULPHYDRIC ACID PRECIPITATE IN AMMONIC SULPHIDE.

(*79*) Add ammonic sulphide to a small quantity of the precipitated sulphides (*76*), and warm gently.

If the precipitate dissolves entirely, it consists of the sulphides of metals of Group VI. Those of Group V.

are absent. Test the remainder of the precipitate according to (**82**) (*a*).

(**80**) If a part of the precipitate does not dissolve, add four or five parts of water, and separate the solution by filtration from the undissolved precipitate.

The part of the precipitate which is insoluble in ammonic sulphide, after being carefully washed, must be tested for sulphides of metals of Group V. (See page 112, IX.)

(**81**) The ammonic sulphide solution obtained in (**80**) may contain metals of Group VI. Add to it gradually dilute chlorhydric acid until the solution becomes acid, and observe the color and general appearance of the precipitate which is produced. It is well to boil the liquid after the formation of a precipitate.

If only a fine white precipitate forms, which remains a long time in suspension in the liquid, even after boiling, it consists of sulphur, and metats of Group VI. are absent, and the tests described in VIII. can be omitted.

A flocculent precipitate, or one that becomes so on boiling, indicates the presence of metals of Group VI., and the color of the precipitate shows what metals predominate. Pass to the following tests:

VIII.—SEPARATION OF METALS OF GROUP VI.

(**82**) If the test (page 108, VII.) has shown the presence of metals of Group VI., and if the precipitate with sulphydric acid (**77**) was not entirely soluble in ammonic sulphide, the whole of that precipitate must be treated two or three times with ammonic sulphide, as directed (**79**), (**80**), and (**81**); and the sulphides of Group VI. must be precipitated from the solution, and, after careful washing, treated as described in (**83**).

(**a**) Were the sulphides, precipitated by sulphydric acid, wholly soluble in ammonic sulphide (see ***79***), it is sufficient to wash and dry the portion of the precipitate (77) which was not treated with ammonic sulphide, and to use it for (***83***).*

(***83***) Free the precipitate (***82***) as completely as possible from water by pressing the filter and its contents between several thicknesses of filter-paper, remove the precipitate from the filter and heat it with concentrated chlorhydric acid. The sulphides of antimony and tin will be dissolved, while the sulphide of arsenic remains undissolved. Collect the residue in a filter, wash, and dry it at 100°, and test for arsenic as follows:

Arsenic. (***84***) Place a small portion of the dried residue in a small tube closed at one end, and put over it about six times its bulk of a mixture of equal parts of sodic carbonate and potassic cyanide. Heat the portion of the tube above this mixture, and afterwards the mixture itself, gently; if as the result of this any moisture is deposited in the upper part of the tube, wipe it out carefully with a rolled-up strip of filter-paper. When the whole is thoroughly dry, heat the lower part of the tube with its contents to a red heat. ARSENIC, if present, is sublimed and deposited on a black or brownish ring in the upper and cooler part of the tube.

Antimony. (***85***) Concentrate the solution of the other two sulphides that was filtered from the arsenious sulphide

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* By heating the dry precipitate in a glass tube, or with less accuracy, by heating it before it is dry, on a bit of glass or porcelain, an approximate test may be made (see ***13***); and in case arsenic sulphide alone is indicated by the complete volatility of the precipitate, this test is conclusive, and the remaining tests in the separation of metals of Group V. may be omitted. The test has little value except when the pure *yellow* color of the precipitate gives rise to the suspicion that only arsenic sulphide is present.

(*83*) in an evaporator, put a small piece of zinc in the concentrated solution, and bring the edge of a piece of platinum foil in contact with the zinc for a minute or two. If ANTIMONY is present the portion of the platinum immersed in the liquid will be stained black by a thin deposit of that metal.

Tin. (*86*) Put the rest of the concentrated solution of the sulphides in an evaporator with more zinc, collect the precipitated black flakes that may appear after a time on a filter, wash carefully, and pour a little concentrated chlorhydric acid on the filter. TIN, if present, will be dissolved, and the solution that passes through will give a white precipitate, or perhaps a gray one if much metal is present, with mercuric chloride. (See page 56.)

Gold. (*87*) (The analytical chemist usually knows whether it is necessary to test for gold or not. In ordinary analyses its presence would be improbable.) When gold is present, it will be found accompanying the sulphide of arsenic remaining after treatment with hot chlorhydric acid (*83*). Heat a portion of this residue in a porcelain crucible or on a piece of a broken evaporator until the arsenic and excess of sulphur have been volatilized. Dissolve the portion which remains in a mixture of chlorhydric and nitric acids. Evaporate the solution nearly to dryness, dilute with water, and add ferrous sulphate solution. The formation of a *brown* or *purple* precipitate of METALLIC GOLD, either immediately or after heating, indicates the presence of the metal.

IX.—SEPARATION OF METALS OF GROUP V.

(SECTION II.)

(*88*) If a portion or the whole of the precipitate obtained with sulphydric acid is insoluble in ammonic sulphide see (*80*), free it by careful washing from the liquid in which it was formed, or from the ammonic sulphide which was used to dissolve the soluble portion, place it in a porcelain dish, pour upon it pure concentrated nitric acid, and heat it gently, if red fumes are given off, until they cease. In any case, complete the operation by adding a little water, and boiling the contents of the dish for a few minutes.

If no part of the precipitate, or if only yellow particles of sulphur remain insoluble, mercury is absent. Pass to (*90*).

In this test, when the liquid holding sulphur in suspension is boiled, the sulphur melts, and may enclose particles of *black* sulphide, which then become very difficult to dissolve, and the appearance of the sulphur may, in such a case, lead to an erroneous conclusion that mercuric sulphide is present. It is for this reason that the precipitate is oxidized with strong nitric acid, as far as possible at a temperature below its boiling point, before the sulphides are finally boiled with a somewhat weaker acid. The same cause makes the confirmatory test for mercury with stannous chloride (*89*) necessary, when there appears to be an insoluble *black* sulphide.

Mercury. If a *black* sulphide, HgS, remains insoluble (*89*) after the above treatment (*88*), a MERCURIC SALT is probably present. Confirmatory test: Boil the *black* insoluble sulphide with chlorhydric acid and a little potassic chlorate in a porcelain dish, and evaporate until the greater part of the acid is volatilized; dilute with water (it is not necessary to filter), and add stannous chloride. A *white* precipitate of MERCUROUS CHLORIDE, Hg_2Cl_2, is formed, if mercury is present.

If mercury is present, dilute a few drops of the nitric acid solution of the sulphides (*88*) with water, and add sulphydric acid. If a black or brown precipitate is formed the solution must be tested for LEAD, BISMUTH, COPPER, and CADMIUM. Pass to (*90*). If the precipitate is yellow, cadmium alone is present (*93*). If no precipitate is obtained all of these four metals are absent. Pass to page 115, X.

Lead. (*90*) Add a few drops of strong sulphuric acid to a small portion of the nitric acid solution of the sulphides (*88*) and evaporate until *dense*, *white* fumes of sulphuric acid appear, and dilute with a considerable quantity of water. If a *white* precipitate forms, it consists of SULPHATE OF LEAD, $PbSO_4$.

The test can be made more delicate by adding an equal bulk of alcohol to the solution after it has been diluted with water. If lead is discovered, treat the whole of the nitric acid solution of the sulphides in the same way; filter and use the filtrate for (*91*).

Bismuth. (*91*) Add ammonic hydrate to alkaline reaction, to the nitric acid solution of the sulphides, or to the filtrate from the lead precipitate, if lead was present. If bismuth is present it is precipitated as the HYDRATE OF BISMUTH, $Bi(HO)_3$, *white*. If bismuth is present, filter and use the filtrate for (*92*) and (*93*).

Copper. (*92*) If copper is present it is dissolved by the ammonic hydrate, and imparts a *blue* color to the solution.

Cadmium. (*93*) If the ammoniacal solution obtained in (*91*) was colorless, copper is absent, but cadmium may be present. In this case add sulphydric acid. A *yellow* precipitate, CdS, indicates CADMIUM.

If copper is present, neutralize the blue ammoniacal solution with chlorhydric acid and add sulphydric acid. This will precipitate both CdS and CuS; the former is soluble in hot dilute sulphuric acid. After washing the mixed sulphides treat the mass with hot dilute sulphuric acid and filter. Cool the filtrate and pass sulphydric acid again through the liquid. If *cadmium* is present there will be a *yellow precipitate* of cadmic sulphide, CdS.

METALS OF GROUPS IV. AND III.

X.—AMMONIC SULPHIDE TEST.

(94) To a small portion of the filtrate from the precipitate produced by sulphydric acid, or to a portion of the original solution, if no precipitate is produced in it by sulphydric acid, add sufficient ammonic hydrate (free from carbonate) to make the reaction alkaline, and then, whether a precipitate is formed or not, add ammonic sulphide. If the solution contains no chlorhydric acid, it is necessary to add a small quantity before neutralizing with ammonic hydrate.

If a *black* precipitate is formed, it may contain the SULPHIDES OF NICKEL, NiS ; COBALT, CoS ; IRON, FeS ; MANGANESE, MnS ; and ZINC, ZnS ; and the HYDRATES OF ALUMINIUM, $Al_2(HO)_6$, and CHROMIUM, $Cr_2(HO)_6$. If the precipitate is *white*, *flesh-colored*, or *light green*, it can only consist of the SULPHIDES OF MANGANESE AND ZINC, and the HYDRATES OF ALUMINIUM AND CHROMIUM. *In this latter case omit* **(97)**, **(98)**, and **(101)**.

If no precipitate is formed no members of Groups III. and IV. are present. Pass to page 123, XII.

(95) If a precipitate was formed in the above test **(94)**, treat a considerable quantity of the solution in the same way, heat the liquid, filter as quickly as possible, and wash immediately with boiling water, until the filtrate has no longer an alkaline reaction.

The filtrate must be tested according to page 123, XII.

SULPHIDES INSOLUBLE IN DILUTE CHLORHYDRIC ACID.

(*96*) Add to the precipitate (***95***) cold, dilute chlorhydric acid; if a *black* residue is insoluble, it consists of the SULPHIDE OF NICKEL OR COBALT. Filter and examine the residue on the filter according to (***97***) and (***98***).

The filtrate must be tested according to (***99***).

If the precipitate dissolves entirely, or if only a white residue is left, no NICKEL *or* COBALT *are present. Pass to* (***99***).

Cobalt. **(*97*)** Dissolve a portion of the precipitate, which proved to be insoluble in cold, dilute chlorhydric acid, in the borax bead, and expose the bead to the action of the outer blowpipe flame. If the bead is *blue*, COBALT is present. If the bead is *brown*, NICKEL is present in large quantity. The bead is *blue* even when more nickel than cobalt is present. To test for traces of cobalt in a nickel bead, detach the hot bead from the wire, heat it two or three minutes on charcoal in a good reducing flame, remove it from the charcoal, and melt it on the platinum wire in the reducing flame. Even if only *traces* of cobalt are present, the bead will be colored blue.

Nickel. **(*98*)** When nickel is present, and when it is nearly free from cobalt, it can be discovered by the *brown* color which it imparts to the borax bead, and the test is conclusive.

To discover small quantities of nickel in the presence of considerable quantities of cobalt, dissolve the precipitate, insoluble in cold dilute chlorhydric acid (see ***96***), in concentrated nitric acid, and neutralize with sodic carbonate. Potassic cyanide is added till the resulting precipitate has dissolved, and then sodic hypochlorite till the liquid smells strongly of it, even after being shaken. It is then boiled. If nickel is present, a *black* precipitate of the sesquioxide (Ni_2O_3) is obtained.

SULPHIDES SOLUBLE IN DILUTE CHLORHYDRIC ACID.

(99) Boil in an evaporating dish the solution which was obtained by the treatment of the ammonic sulphide precipitate with cold, dilute chlorhydric acid (see **96**), until the smell of sulphydric acid has entirely disappeared ; add a few drops of strong nitric acid, and boil a minute longer, and filter if a precipitate of sulphur has formed.

If oxalic, phosphoric, and boracic acids are not known to be absent, pass to page 119, XI.

(100) Add sodic hydrate to the solution obtained in (**99**) until the reaction becomes very strongly alkaline, dilute with water and boil for a few minutes.

If a precipitate forms immediately, or after boiling, test it according to (**101**) and (**102**) ; it may contain the HYDRATES of IRON, $Fe_2(HO)_6$, MANGANESE, $Mn(HO)_2$, and CHROMIUM, $Cr_2(HO)_6$. In this case filter and test the filtrate, which may contain ZINC and ALUMINIUM, according to (**103**) and (**104**).

If no precipitate forms, pass to (**103**). The solution may contain ZINC and ALUMINIUM.

Iron. **(101)** Put a small quantity of the precipitate (**100**) on a watch-glass, dissolve it in a single drop of dilute chlorhydric acid, dilute with water, and add potassic sulphocyanate. A *red* color indicates the presence of IRON.

Manganese and ***Chromium.*** **(102)** Fuse a portion of the precipitate obtained in (**100**) on platinum foil with sodic carbonate and sodic or potassic nitrate. A *green* mass indicates MANGANESE, owing to the formation of sodic manganate ; a *yellow* mass, CHROMIUM, from the formation of the sodic chromate. A very small quantity of manganese can readily be detected in the presence of a considerable quantity of chromium by the green color imparted to the mass. In case of doubt, however, place the platinum foil in an evaporator, cover it with water, add a few drops of alcohol, and boil. The sodic chromate will be entirely dissolved, along with the excess of sodic carbonate, while the sodic manganate will be decomposed, yielding brown flakes of manganese sesquioxide. These can be filtered through a small filter, washed thoroughly with hot water, and fused again upon platinum foil, as before, when the green color will appear very distinctly. To detect a small quantity of chromium in presence of excess of manganese, treat the mass as above with boiling water and filter. The appearance of a yellow color in the solution is proof of the presence of CHROMIUM.

Acetate of lead may also be added to the solution, acidified with acetic acid, and it then takes a deeper *yellow* color and becomes turbid, or a *yellow* precipitate is formed if chromium is present. The formation of a white precipitate under these circumstances only indicates that the carbonate of sodium employed contained sulphate of sodium as an impurity.

Zinc. ***(103)*** Add sulphydric acid to a portion of the solution in sodic hydrate, obtained in (***100***), after it has been filtered, if a precipitate was formed. If a *white*, *flocculent* precipitate forms, it consists of SULPHIDE OF ZINC, ZnS.

If chromium was discovered in (***102***), zinc may also be present in the precipitate obtained in (***100***). Therefore, in that case, dissolve a portion of the precipitate by boiling with a very little dilute chlorhydric acid, add sodic hydrate until the reaction is alkaline, acidify with acetic acid and add sulphydric acid. A *white, flocculent* precipitate consists of SULPHIDE OF ZINC, ZnS.

Aluminium. **(*104*)** Add to another portion of the sodic hydrate solution, obtained in (***100***), chlorhydric acid until the reaction becomes acid, and then ammonic hydrate until it becomes alkaline, and boil. If a *white, flocculent* precipitate forms, it consists of ALUMINIC HYDRATE, $Al_2(HO)_6$. This precipitate is at first gelatinous, and it may easily escape notice; it is therefore best to set the test-tube aside, and to wait a quarter of an hour for the precipitate to settle.

XI.—AMMONIC SULPHIDE TEST IN CASE PHOSPHORIC, OXALIC, AND BORACIC ACIDS ARE PRESENT.

If PHOSPHORIC, OXALIC, or BORACIC ACID was present in the original solution, these acids, together with the metals of Group II., may be contained, wholly or in part, in the solution obtained in (***99***), for the acids would be precipitated with any of the metals of Groups II., III., and IV., during the treatment with ammonic hydrate and ammonic sulphide in (***94***), and the precipitates would be dissolved during the treatment with dilute chlorhydric acid in (***96***), and consequently the metals and acids might be contained in the solution (***99***). The solution (***99***) must be tested for phosphoric, oxalic, and boracic acids, and freed from them before the ordinary course of analysis can be proceeded with.

Phosphoric Acid. Use the test (***127***) with a small portion of the solution obtained in (***99***).

If phosphoric acid is present, use (**107**) *and the succeeding tests.*

If phosphoric acid is absent, use (**100**) *and the succeeding tests.*

In either case first perform the operation described in the next paragraph.

(**105**) Before testing for OXALIC AND BORACIC ACIDS it is necessary to set them free from their combinations with the metals of Groups II., III., and IV. To effect this end add sodic carbonate to a small quantity of the solution obtained in (**99**) until the reaction becomes strongly alkaline, and boil for a few minutes and filter. The metals are precipitated, with the exception of a portion of the aluminium, and the acids remain in the solution.

[If no precipitate is formed with sodic carbonate it is unnecessary to test further for these metals or acids, as in that case they cannot be present in the solution (**99**).]

Oxalic Acid. Test a portion of the filtrate obtained after boiling with sodic carbonate for oxalic acid according to (**125**).

Boracic Acid. Test another portion of the same filtrate for boracic acid according to (**128**).

If one or both these acids are found, the whole of the remainder of the solution obtained in (**99**) *must be treated with sodic carbonate as in* (**105**), *and the precipitate thus obtained must be dissolved in dilute chlorhydric acid. The solution must be used for* (**107**) *and the succeeding tests, if phosphoric acid was discovered.*

If phosphoric acid is absent, the solution must be used for (**100**) *and the succeeding tests.*

Aluminium. If it was found necessary to treat the solu-
(**106**) tion obtained in (**99**) with sodic carbonate according to (**105**), the filtrate from the precipitate produced

by sodic carbonate may contain a portion of the aluminium. Add to the filtrate in (***105***) chlorhydric acid until the reaction becomes acid, and then ammonic hydrate until the reaction becomes alkaline. If a *white, flocculent* precipitate forms immediately, or after long standing, it contains aluminium.

Iron. (***107***) Add a few drops of potassic sulphocyanate to the solution obtained in (***99***) ; a *red* color indicates the presence of IRON.

(***108***) Add to the remainder of the solution obtained in (***99***), if phosphoric acid alone is present, or to the solution obtained in (***105***), if oxalic or boracic acid is likewise present, ferric chloride, until a few drops, treated with ammonic hydrate on a watch-glass, give a *yellow* and not a *white* precipitate ; dilute largely with water, render the solution alkaline with ammonic hydrate, and add a considerable excess ; then add acetic acid until the solution has a slight acid reaction, and boil for a few minutes in a flask.

The precipitate contains all the IRON, ALUMINIUM, CHROMIUM, and PHOSPHORIC ACID present in the solution so treated. Test according to (***109***) and (***110***).

The filtrate contains the MANGANESE, ZINC, and probably part of the BARIUM, CALCIUM, and MAGNESIUM which were present in the original solution.

The operations described, page 115, X., *and* page 123, XII., *must be repeated with this filtrate, omitting those which relate to the separation of* NICKEL *and* COBALT, *and the detection of* IRON, CHROMIUM, *and* ALUMINIUM.

Chromium. (***109***) Test a portion of the precipitate obtained in (***108***) for chromium according to (***102***).

Aluminium. (***110***) Boil the remainder of the precipitate obtained in (***108***) with sodic hydrate, and test the solution for aluminium according to (***104***).

[The method of precipitation by boiling the acetic acid solution used in (***108***) can be used in all cases for the separation of aluminium, chromium, and iron (ferric salts) from the metals of all other groups, except the Groups V. and VI., and it is preferable to any other, but it demands more skill in manipulation.]

METALS OF GROUP II.

XII.—DETECTION OF BARIUM, STRONTIUM, CALCIUM, AND MAGNESIUM.

Should the filtrate, after removal of the metals of Groups III. and IV., have a brown color, it can only come from the presence of nickel, a small quantity of the sulphide of that metal having been dissolved in the excess of ammonic hydrate and ammonic sulphide used. Before proceeding to the detection of metals of Group II., the nickel must be entirely removed, and this can be readily accomplished by boiling for a few minutes and filtering again.

(*111*) To a small portion of the solution, to which the previous tests have been applied, or to a solution which has been found to contain no metals of the higher groups, add ammonic hydrate until the reaction becomes alkaline, and then ammonic carbonate, and boil. (If the solution does not already contain ammonic chloride, this also must be added, to prevent the precipitation of magnesium.)

If a white precipitate forms, it can only consist of BARIC CARBONATE, $BaCO_3$, STRONTIUM CARBONATE, $SrCO_3$, and CALCIC CARBONATE, $CaCO_3$.

If no precipitate forms, BARIUM, STRONTIUM, *and* CALCIUM *are absent; pass to* **(*116*)**.

If a precipitate was formed with ammonic carbonate, the whole of the solution must be treated as described.

above. Filter, wash the precipitate, and test the filtrate according to (*116*). Dissolve the precipitate by pouring a very little dilute chlorhydric acid on the filter, and use the solution thus obtained for (*112*), (*113*), (*114*), and (*115*).

Barium. (*112*) To a small portion of the solution in chlorhydric acid add a considerable quantity of calcic sulphate. If a precipitate forms immediately, it consists of BARIC SULPHATE, $BaSO_4$.

Strontium. (*113*) If on the addition of calcic sulphate, as in (*112*), a precipitate appears only after some little time, it consists of STRONTIUM SULPHATE, $SrSO_4$.

(*114*) If BARIUM OR STRONTIUM is discovered by means of calcic sulphate, add dilute sulphuric acid to another portion of the chlorhydric acid solution. Boil, filter, and test the filtrate for calcium (*115*). Examine the precipitate on platinum wire moistened with chlorhydric acid in the flame. (See pages 26 and 34.) Barium and strontium can both be detected, even when a small quantity of one is present with a large quantity of the other. After placing a small particle of the precipitate on the loop of the platinum wire the particle should be repeatedly moistened in chlorhydric acid and subjected again to the action of the flame. Where strontium is present in very small proportion the barium color will, after repeated moistening with chlorhydric acid, finally give place to the crimson of strontium. Where, however, the reverse proportion is found, the detection of barium is not so easy by this method. In this case add to a portion of the chlorhydric acid solution obtained in (*111*) a solution of strontium sulphate. A faint white cloudiness appearing after some time indicates the presence of barium.

Calcium. (*115*) If barium, or strontium, or both, were present in the precipitate obtained on the addition of ammonic carbonate, add ammonic hydrate to alkaline reaction to the filtrate obtained after precipitation with dilute sulphuric acid (***114***), and then ammonic oxalate. If a *white* precipitate forms, it consists of CALCIC OXALATE, CaC_2O_4.

If neither barium nor strontium was found, the whole of the precipitated carbonate (***111***) must have consisted of calcic carbonate. Confirm by flame reaction, page 35.

It is quite possible that owing to the presence of a large excess of ammonic chloride in the solution, the calcium may have escaped precipitation by ammonic carbonate. In this case it would come down as a flocculent precipitate in the test for magnesium (***116***). To detect calcium under these circumstances filter the precipitate obtained in (***116***) and dissolve it in acetic acid, dilute considerably and add ammonic oxalate. If calcium was contained in the precipitate obtained by hydric disodic phosphate, it will now be precipitated as the oxalate, which is insoluble in acetic acid. Filter the calcic oxalate, and to the filtrate add ammonic hydrate to alkaline reaction, when the ammonio-magnesic phosphate will be re-precipitated.

Magnesium. **(*116*)** To the filtrate from the precipitate produced by ammonic carbonate (***111***), or to the solution in which no precipitate was obtained on addition of that reagent, add sodic phosphate. If a *white* precipitate forms (frequently only after the lapse of some minutes), it consists of the AMMONIO-MAGNESIC PHOSPHATE, $MgNH_4PO_4$.

METALS OF GROUP I.

XIII.—DETECTION OF SODIUM, POTASSIUM, AND AMMONIUM.

Before testing for sodium and potassium, precipitate the metals of Groups V. and VI. which are present with sulphydric acid, and precipitate those of Groups II., III., and IV. with a mixture of ammonic carbonate and sulphide, and use the solution, freed from those metals, for the following tests (***117***) and (***118***):

Sodium. Evaporate the solution to dryness, and drive off (***117***) ammonia salts, if any are present, by heat. Sodium, if present, can be distinguished by the *yellow* color which a small quantity of the solid residue held in the flame of a gas or alcohol lamp imparts to it. See page 26.

Frequently sodium can be detected in a solution without evaporation, by dipping a platinum wire in the solution, and then holding it in the flame.

Potassium can be recognized by the violet color which it (***118***) imparts to the flame. The solid residue obtained in (***117***) can be tested for potassium in the same way that it is tested for sodium.

If sodium is also present, its greater coloring power will obscure the potassium flame, but by looking through a piece of blue glass at the flame, the violet color can be distinguished even when sodium is present. (The color of the potassium flame is almost the same as that of the heated wire, while the sodium flame is much more blue, if it is not excluded entirely by the glass.)

Ammonium. Add to a portion of the original solution a few drops of sodic hydrate, and heat.
(*119*)

AMMONIA can be recognized by its smell, or by holding a piece of moistened turmeric paper or red litmus paper at the mouth of the test-tube, taking care not to let it touch the sides, which may be moistened with sodic hydrate. The turmeric paper will be turned brown, and the litmus paper blue, if ammonia is present.

TESTS FOR ACIDS.

It is usual to take a fresh quantity of the solution to test for acids, and generally the tests for metals precede those for acids, in order that information gained by the first series of experiments may point out the most convenient way of detecting the acids. Silicic acid, however, is always first precipitated from a solution as directed, page 101 (*64*), and the acids noticed under the chlorhydric acid test (page 106, V.) are to be looked for while that test is applied to the detection of the metals. Phosphoric, oxalic, and boracic acids interfere with the ordinary methods of testing for the metals of Groups II., III., and IV., and consequently these acids must be tested for as directed, page 119, XI., during the application of the tests for the metals of those groups.

In all other cases, when the presence of any metals or acids in the solution interferes with the performance of the test for an acid, directions are given under the head of each acid for removing them.

It is obvious that metals and acids which precipitate each other cannot be present together in a solution, and that when certain metals have been found, the number of acids which it becomes necessary to look for is restricted within limits determined by this consideration. Therefore the knowledge already acquired of the composition of a solution must be brought to bear upon the problem of testing for acids.

First, the reaction which the solution gives with test-paper must be considered, and then the tables IV. and V. must be consulted to ascertain what acids can exist in a solution possessing the observed reaction, together with the metals which have been discovered.

For instance, if a solution contains *lead*, and is neutral or nearly neutral, the only acids which can be present in it are acetic, chlorhydric, chloric, and nitric. If the reaction is strongly acid the solution may contain all the acids except sulphuric, sulphydric, ferro- and ferricyanhydric. Moreover, the solution cannot contain a large quantity of lead and chlorhydric acid at the same time, because the chloride of lead is only soluble in 135 parts of water.

ACIDS OF GROUP I.

Arsenious and Arsenic Acids are always detected (***120***) by the sulphydric acid test in searching for the metals. When these acids are discovered, they must always be precipitated by sulphydric acid before testing further.

Chromic Acid cannot be present in a solution to which (***121***) SULPHYDRIC ACID or AMMONIC SULPHIDE has been added (see page 64). If chromic acid is present in a solution, it must be contained in the precipitate obtained with baric chloride (***122***), and can be detected by heating a small portion of the precipitate in the borax bead. If CHROMIC ACID, H_2CrO_4, is present, the bead will be colored *green*. Chromic acid can often be recognized by the *yellow* color which it imparts to solutions which contain it, and by the yellow precipitate, $PbCrO_4$, which is obtained by adding plumbic acetate to the neutral or slightly acid solution.

XIV.—BARIC CHLORIDE TEST.

Baric nitrate and nitric acid should be used instead of baric chloride and chlorhydric acid, when lead, silver, or mercurous salts have been discovered in the solution.

(122) Put a piece of litmus paper in the solution, and if the reaction is acid, add ammonic hydrate, drop by drop, until it becomes slightly alkaline. If a precipitate is formed in consequence, add dilute chlorhydric acid only in sufficient quantity to dissolve it. Add baric chloride, and if a precipitate forms, it indicates the presence of ARSENIOUS, ARSENIC, CHROMIC, SULPHURIC, SULPHUROUS, OXALIC, FLUORHYDRIC, PHOSPHORIC, BORACIC, CARBONIC, OR SILICIC ACIDS. (Silicic acid cannot be present after the operation (***64***) has been performed).

If these acids are absent, pass to the argentic nitrate test (page 132, XVIII.).

(The baric chloride test is of little value except when the substance is soluble in water with a neutral or slightly alkaline reaction, or in case the reaction is acid, when the metals of Groups II., III., IV., and V. are absent. The following special tests are more accurate :)

Sulphuric Acid. ***(123)*** Acidify (if the solution is not already acid) with dilute chlorhydric acid, in considerable excess, and add baric chloride as long as a precipitate continues to form. (Use nitric acid and baric nitrate if chlorhydric acid produces a precipitate.) If sulphuric acid is present, it is precipitated as BARIC SULPHATE, $BaSO_4$, *fine white powder.* The solution must not be heated when it is intended to use the filtrate for the next test.

Sulphurous Acid. ***(124)*** To the filtrate from the precipitate produced by baric chloride, or to the acid solution to which baric chloride has been added without producing a precipitate, add potassic dichromate, and boil. If a precipitate forms, it consists of BARIC SULPHATE, $BaSO_4$, produced by the oxidation of SULPHUROUS ACID, H_2SO_3, contained in the solution. Usually the CHLORHYDRIC ACID TEST (***75***) is more convenient and sufficiently accurate.

XV.—CALCIC SULPHATE TEST.

If the solution contains metals which are precipitated by sulphydric or sulphuric acids, they must first be removed by adding a slight excess of those precipitants and filtering.

Oxalic Acid, (***125***) if the solution is alkaline, add acetic acid until the reaction becomes acid; if it is acid, add sodic hydrate until the reaction becomes alkaline, and then add acetic acid until it becomes acid, and test as below for oxalic acid. If a precipitate forms and does not dissolve in the acetic acid, add to the original solution a considerable excess of sodic carbonate, boil, filter, add to the filtrate acetic acid until its reaction becomes acid, and test as follows for oxalic acid: Add calcic sulphate in considerable quantity. If a precipitate forms, it consists of CALCIC OXALATE, CaC_2O_4, *white powder.* Fluorhydric acid is the only other acid which precipitates calcic sulphate under these circumstances, and as the precipitate is *almost transparent* and *gelatinous*, it cannot easily be mistaken for that produced by oxalic acid.

Fluorhydric Acid (***126***) can only be present in alkaline solutions in glass vessels. If there is reason to suspect the presence of this acid, add calcic chloride and ammonic hydrate to the solution, and if a precipitate forms, collect it on a filter, and examine it for fluorine according to (***44***).

XVI.—AMMONIC MOLYBDATE TEST.

Phosphoric Acid. Make the solution strongly acid (if (**127**) it is not so already) with nitric acid, and add a small portion of it to a considerable quantity of ammonic molybdate solution.* If phosphoric acid is present, PHOSPHO-MOLYBDATE OF AMMONIUM, *yellow crystalline powder*, is precipitated. If the quantity of phosphoric acid in the solution is very small, the precipitate does not form until after several hours.

If sulphydric acid is present, it is necessary to heat the acid solution until it is expelled, before performing the test. See (**136**) (*a*) for this test in the presence of ferro- or ferricyanhydric acid.

XVII.—TURMERIC PAPER TEST.

Boracic Acid.† Strongly acidify the solution (if it is not (**128**) already acid) with dilute chlorhydric acid, dip a piece of turmeric paper in it, and dry the paper by holding it over the lamp-flame, without charring it. If a *red* or *brownish-red* stain appears upon the paper when it is dry, it is due to the presence of BORACIC ACID, H_3BO_3.

ACIDS OF GROUP II.

XVIII.—ARGENTIC NITRATE TEST.

Acidify with nitric acid, if the solution is not already acid, and add argentic nitrate.

Sulphydric Acid. If a *black* precipitate is formed, it (**129**) must contain ARGENTIC SULPHIDE, Ag_2S, showing

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* See foot-note, page 67.

† Ibid., page 68.

that sulphydric acid was present in the solution. The precipitate may also contain ARGENTIC CHLORIDE, BROMIDE, IODIDE, CYANIDE, FERRO- and FERRICYANIDE.

If a precipitate is formed which is not *black*, it can only be due to presence of CHLORHYDRIC, BROMHYDRIC, IODHYDRIC, CYANHYDRIC, FERROCYANHYDRIC, and FERRICYANHYDRIC ACIDS.

If no precipitate is formed, none of the above acids are present. Pass to (***137***).

Chlorhydric Acid. See also (***48***). Acidify strongly (***130***) with concentrated nitric acid, add argentic nitrate in excess, shake thoroughly if a precipitate forms, and allow it to settle, decant the liquid, and pour on the precipitate strong nitric acid, and boil for five minutes; if a precipitate remains undissolved, it consists of ARGENTIC CHLORIDE, AgCl, if violet. If yellow, of ARGENTIC BROMIDE. See (***132***) (*a*).

See (***136***) (*a*) *for this test in the presence of ferro- or ferricyanhydric acid.*

Bromhydric Acid gives with argentic nitrate a white (***131***) curdy precipitate, which darkens on exposure to light, is insoluble in boiling concentrated nitric acid, and is not so readily soluble in ammonia as the argentic chloride. See also (***49***).

Iodhydric Acid gives with argentic nitrate a yellow pre- (***132***) cipitate of the iodide which is nearly insoluble in ammonia. See also (***50***).

(***132***) (*a*) *Detection of* CHLORHYDRIC, BROMHYDRIC, *and* IODHYDRIC *acids in presence of each other.* Mix the liquid to be tested with a few drops of dilute sulphuric acid, then with a little starch paste, and add a few drops of fuming nitric acid, or a solution of hyponitric oxide in sulphuric acid, when the blue color characteristic of iodine will appear. (See page 74.) Add now chlorine water until that reaction has disappeared. On continuing the gradual addition of chlorine water the bromine will be set free, and will impart a yellow or brownish color to the liquid. (See page 73.) Chlorine in presence of bromine can best be detected by the following method : Evaporate the solution to dryness and mix the residue with a little potassic dichromate. Place the mixture at the bottom of a clean test-tube and pour on it a few drops of concentrated sulphuric acid. On the application of heat dark red drops of CHROMYL DICHLORIDE, or CHROMIC OXYCHLORIDE, CrO_2Cl_2, will condense in the upper portion of the test-tube. Bromides, under the same treatment, give a similar result ; but in the latter case the distillate consists of *bromine*, which is instantly decolorized by a drop of ammonic hydrate. In the case of chlorides, when the CrO_2Cl_2 is treated with ammonic hydrate, it gives a yellow solution, owing to the formation of ammonic chromate $[(NH_4)_2CrO_4]$.

Where the substance to be tested for bromine and iodine was not soluble in water it should be fused on platinum foil with sodic carbonate. The mass is then treated with water and the aqueous solution used for the foregoing method of separation.

XIX.—PRUSSIAN BLUE TEST FOR CYANHYDRIC ACID.

Cyanhydric Acid. Add to the solution ferrous sulphate and a few drops of ferric chloride ; add sodic hydrate until a precipitate forms (unless the solution is alkaline and a precipitate forms without the addition of sodic hydrate), warm for a minute, and add dilute chlorhydric acid until the reaction becomes acid.
(***133***)

The appearance of a *blue* precipitate or a *blue* color in the solution is evidence of the presence of cyanhydric acid.

See (**136**) (*b*) *for this test in the presence of ferro- or ferricyanhydric acid.*

XX.—FERRIC CHLORIDE TEST.

Ferrocyanhydric Acid. Add a little ferric chloride (**134**) to the acid solution. If ferrocyanhydric acid is present, a precipitate of PRUSSIAN BLUE, $Fe_4(FeCy_6)_3$, *deep blue*, is formed.

XXI.—FERROUS SULPHATE TEST.

Ferricyanhydric Acid. Add a little ferrous sulphate (**135**) to the acid solution. If ferricyanhydric acid is present, a precipitate of TURNBULL'S BLUE, $Fe_3(Fe_2Cy_{12})$, *deep blue*, is formed.

(**136**) (*a*) If ferro- or ferricyanhydric acid is present, before performing the tests for PHOSPHORIC ACID (**127**), and CHLORHYDRIC ACID (**130**), the following steps must be taken: Add dilute sulphuric acid, dilute with water if the solution is not dilute, add cupric sulphate, and finally add enough baric nitrate* to render the precipitate of a decidedly lighter color; heat almost to boiling, allow the precipitate to settle for a few minutes, filter, and use the filtrate for the tests (**127**) and (**130**). In case the test (**51**) for nitric acid is to be used take the same preliminary steps; using baric chloride in place of baric nitrate.

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* Sulphuric acid and baric nitrate or chloride are only added in order to produce a heavy precipitate of baric sulphate, which carries down with it the lighter particles of the other precipitates, and renders the filtration easier.

(136) **(b)** If ferro- or ferricyanhydric acid is present, the test for CYANHYDRIC ACID **(133)** is to be modified in the following manner: Dilute with water if the solution is not dilute, add dilute sulphuric acid, then add, according as ferro- or ferricyanhydric acid is present, ferric chloride or ferrous sulphate, or both, in sufficient quantity to precipitate the ferro- or ferricyanhydric acid or both acids; finally, add baric chloride * until the color of the precipitate has become decidedly lighter; shake thoroughly, allow the precipitate to settle for a few minutes, and filter. Add to a portion of the filtrate sodic hydrate until a precipitate forms, warm gently, and add dilute chlorhydric acid until the solution becomes acid. The appearance of a blue precipitate, or of a *blue color*, is evidence of the presence of CYANHYDRIC ACID.

ACIDS OF GROUP III.

ACIDS WHICH ARE NOT PRECIPITATED BY ANY METALS.

Chloric Acid. **(137)** See sulphuric acid test **(47)**.

Nitric Acid. **(138)** See sulphuric acid test **(51)**.

Acetic Acid. **(139)** See sulphuric acid test **(52)**.

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* See note, page 135.

CLASS III.

SUBSTANCES WHICH ARE INSOLUBLE IN WATER AND IN ACIDS.

(See page 97.)

The only substances which are insoluble after the treatment described on page 97 are the following:

Plumbic Sulphate (not absolutely insoluble in acids).

Argentic Chloride (slightly soluble in chlorhydric acid).

Sulphur.

Carbon.

Baric Sulphate, Silica, and many Silicates, and some Oxides.

XXII.—SOLUTION IN AMMONIC ACETATE AND POTASSIC CYANIDE.

Plumbic Sulphate. Boil a portion of the substance (***140***) with ammonic acetate, and test the solution (after filtration, if necessary) with ammonic sulphide. If LEAD is present, a *black* color, or a *black* precipitate is formed.

Test the solution also for SULPHURIC ACID, according to (***123***).

If lead is discovered, repeat the treatment with ammonic acetate until no more lead is dissolved.

Argentic Chloride. Digest a portion of the substance, (***141***) free from plumbic sulphate with potassic cyanide, warm (unless it *blackens* by warming), and test the solution (after filtration, if necessary) with ammonic sulphide. A *black* precipitate indicates the presence of SILVER. If a *black* precipitate forms, wash it, dissolve it in strong nitric acid, and test with chlorhydric acid according to (***68***) in order to confirm the presence of SILVER.

If silver is present, repeat the treatment with potassic cyanide until no more silver is dissolved.

Sulphur. Test the substance, free from plumbic sulphate (***142***) and argentic chloride, for sulphur according to (***10***). If the substance is moist, it must be carefully dried by heating it in a porcelain dish over a water-bath before applying the test.

If sulphur is present, heat the substance in a covered porcelain crucible until the sulphur is completely volatilized.

Carbon. If the substance has a *black* or *gray* color, which it (***143***) loses when it is heated with the blowpipe on platinum foil, carbon in some form is probably present. If carbon is present, the substance, free from plumbic sulphate, argentic chloride, and sulphur, should be burnt, until as much as possible of the carbon is destroyed, by heating it red-hot on platinum foil or in a porcelain crucible.

XXIII.—FUSION WITH POTASSIC AND SODIC CARBONATES AND SODIC NITRATE.

Baric Sulphate, Silicic Acid, and many Sili- (***144***) ***cates, and some Oxides.*** Mix the finely pow- (***a***) dered substance, free from plumbic sulphate, argentic chloride, and sulphur, and as nearly free from carbon as possible, with two parts of potassic carbonate, two parts of sodic carbonate, and one part of sodic nitrate ;* bring as much of the mixture as can be heated at once on the platinum foil, and heat the under side of the foil with a blast-lamp until the whole mass is in a state of quiet fusion. Repeat this operation two or three times, if much substance is required for the analysis.

(*b*) Detach the fused mass from the platinum foil each time by plunging the foil, while it is hot, in distilled water. Boil the product of fusion with water, and if it does not dissolve completely, filter, and wash the precipitate on the filter with distilled water, rejecting the washings. Continue the washing until baric chloride ceases to produce a precipitate in the water which runs through the filter.

(*145*) The first filtrate may contain ARSENIC ACID, see **(*120*)**, (its occurrence is rare) ; CHROMIC ACID, see **(*121*)** ; SULPHURIC ACID **(*123*)**, (the tests referred to above may be applied successively to a single portion of the filtrate) ; FLUORHYDRIC ACID (its occurrence is rare), see **(*126*)** and **(*44*)**, and PHOSPHORIC ACID **(*127*)**. The two last tests may be applied successively to another portion of the filtrate. No compound of these acids, except BARIC SULPHATE, is by itself insoluble, but insoluble substances sometimes contain small quantities of the acids. CALCIC FLUORIDE is only decomposed completely by the treatment with sulphuric acid described in **(*44*)**.

Silicic Acid. The principal portion of the filtrate should **(*146*)** be tested according to **(*64*)** for silicic acid. After

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* The sodic nitrate is added in order to destroy carbon or other reducing substances. If the substance to be analyzed appears to contain much carbon, increase the quantity of sodic nitrate. If the substance contains no carbon, the use of sodic nitrate is usually unnecessary.

separation of silica the only metals* that can be present in the chlorhydric acid solution are LEAD, ALUMINIUM, and ZINC. Test for lead by adding an excess of dilute sulphuric acid and alcohol to the solution. If a precipitate of PLUMBIC SULPHATE forms, filter. Test for ALUMINIUM, in the solution, free from lead, by adding ammonic hydrate in excess. If a precipitate of ALUMINIC HYDRATE forms, filter. Test for ZINC in the solution, free from lead and aluminium by adding to the solution containing ammonic hydrate in excess ammonic sulphide. *A flocculent, white* precipitate consists of SULPHIDE OF ZINC.

(***147***) If a portion remains insoluble after boiling the fused mass with water (***144***) (***b***), dissolve it in chlorhydric acid. If much silica was discovered (see ***146***), it is best to evaporate the chlorhydric acid solution to dryness, and to proceed as directed in (***64***).

Test for metals in the chlorhydric acid solution according to page 107, VI., etc.

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* It is evident that sodium and potassium in insoluble silicates cannot be detected by this process. All reliable methods for their detection require the use of platinum vessels and great care in manipulation. Larger works on analysis must be consulted for such methods.

EXPLANATION OF TABLES.

The Tables I., II., and III. contain a synopsis of the course of analysis of bodies in solution given in Part III., and they are intended as an index to the methods which are there described in detail. They may also serve as guides in analytical work to students who have made themselves acquainted with the detailed descriptions of Part III.

A skeleton form, similar to that of the tables, should be filled out with the results of an analysis, and the reactions which occur on the application of each test should be noted.

The sign —⌣— placed under the formula of a compound indicates that it is formed as a precipitate during a reaction. This sign is used in the following tables, and it will also be found convenient in noting the results of analyses.

The Tables IV. and V. are intended to indicate the degree of solubility in water, and in many cases in alcohol, acids, and alkalies, of the combinations of the metals and acids mentioned in Part II.

The properties of a salt are described in the square formed by the intersection of the column devoted to an acid with that devoted to a metal.

The Roman numerals standing after the symbols of the metals indicate their quantivalence, and the formula of a salt is made by putting the symbol of a metal in the place of the symbol of an equivalent number of atoms of hydrogen in an acid. When an acid contains more than one atom of hydrogen, several classes of salts may be formed, according as one or more atoms of hydrogen are replaced by a metal. The normal or regular salts are those which are formed by the replacement of the greatest possible number of atoms of hydrogen by a metal.

The descriptions of the tables refer to normal salts, but the following cases are exceptions, because the salts specified are more commonly met with in analysis ; and in using the tables the formulas below must be substituted for those of the normal salts :

ARSENATES.—$MgNH_4AsO_4$; $MnNH_4AsO_4$; $(Hg_2)HAsO_4$; $HgHAsO_4$.

PHOSPHATES.—$(NH_4)_2HPO_4$; $BaHPO_4$; $CaHPO_4$; $MgNH_4PO_4$; $MnNH_4PO_4$; $HgHPO_4$; Na_2HPO_4.

The ARSENATE OF ALUMINIUM probably contains more acid than the normal salt.

The CHROMATES OF ALUMINIUM and of IRON (ferric chromate) contain a larger proportion of metal than the normal salts.

ARSENITES.—The arsenites referred to in the table have only *two* atoms of hydrogen replaced by a metal, except $Mg_3(AsO_3)_2$ and Ag_3AsO_3.

The ARSENITES OF COBALT and MANGANESE contain less than two atoms of hydrogen replaced by the metal.

BORATES.—$(NH_4)_2B_4O_7$; BaB_2O_4 ; CuB_4O_7 ; FeB_4O_7 ; $(Fe_2)B_3O_6$; PbB_4O_7 ; CaB_2O_4 ; MnB_4O_7 ; NiB_4O_7 ; $K_2B_2O_4$; $Ag_2B_2O_4$; $Na_2B_4O_7$; ZnB_4O_7.

EXPLANATION OF SIGNS IN TABLES IV. AND V.

Numbers = number of parts of water required to dissolve one part of the anhydrous salt[*] at the ordinary temperature.

∞. = insolubility.[†] The sign of infinity indicates that an infinite quantity of water is required to dissolve the salt.

s. = soluble to a considerable extent in water.

s.s. = slightly soluble.

del. = deliquescent, or capable of dissolving by attracting moisture from the air.

dec. = decomposed. *Examples:* dec. = decomposed by water. — dec. = decomposed by acids.

— = acids. *Example:* — s. = soluble in acids.

+ = sodic or potassic hydrate. *Example:* + s. = soluble in sodic or potassic hydrate.

am. = ammonic hydrate.

am. cl. = ammonic chloride.

al. = alcohol. *Example:* al. ∞. = insoluble in alcohol.

When no solvent, such as —, am., al., etc., is indicated, the signs ∞., s., s.s., and dec. refer to the action of water on the salt.

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[*] The salt without water of crystallization is referred to.

[†] Most of the salts marked insoluble in the table are not really more insoluble than salts like baric sulphate, but they are described as insoluble because they are known to be nearly so, and because the quantity of water required to dissolve them has not been determined.

TABLE I.—*Tests for Metals, to be used successively.*

CHLORHYDRIC ACID TEST.

Metals of Group V., Section I., in acid solution : p. 104, III.

Lead. (67) $\underbrace{PbCl_2}$, *white, crystalline* ; sol. in boiling water. The solution, $+ H_2SO_4 +$ alcohol, gives $\underbrace{PbSO_4}$, *white, pulverulent.*

Silver. (68) $\underbrace{AgCl}$, *white, curdy* ; sol. in NH_4HO. The solution, $+ HNO_3$ to acid reaction, gives $\underbrace{AgCl}$.

Mercurous Salts. (69) $\underbrace{Hg_2Cl_2}$, *white, pulverulent* ; insol. and turns black with NH_4HO.

Metals in alkaline solution : p. 105, IV.

$\underbrace{PbCl_2}$; $\underbrace{PbSO_4}$; $\underbrace{AgCl}$; when the precipitate is white. See (71) (a) See (71) (*a*).

$\underbrace{As_2S_3}$; $\underbrace{Sb_2S_3}$; $\underbrace{Sb_2S_5}$; $\underbrace{SnS}$; $\underbrace{SnS_2}$; where the precipitate is *colored*. See (71) (*b*). (71) (b).

Acids, usually in alkaline solution : p. 106, V.

Silicic Acid, (64) $\underbrace{SiO_2}$, insol. in acids after drying. See p. 71,

Carbonic Acid, (72) CO_2. The gas gives $\underbrace{CaCO_3}$ in a drop of lime-water.

Cyanhydric Acid, (73) HCN. Smell of bitter almonds.

Sulphydric Acid, (74) H_2S. Smell of rotten eggs. The gas blackens lead-paper.

Sulphurous Acid, (75) SO_2. Smell of burning sulphur. The gas turns *blue* a mixture of ferric chloride and potassic ferricyanide.

SULPHYDRIC ACID TEST : p. 107, VI.

Metals of Groups V. and VI. in acid solution.

SOLUBLE IN AMMONIC SULPHIDE : p. 108, VII.

$\underbrace{As_2S_3}$, *yellow* ; $\underbrace{Sb_2S_3}$ and $\underbrace{Sb_2S_5}$, *orange* ; $\underbrace{SnS}$, brown ; $\underbrace{SnS_2}$, *yellow* ; Au_2S_3, *dark brown.*

Precipitate treated with HCl, Au_2S_3, As_2S_3, insoluble, — Sb_2S_3, Sb_2S_5, SnS, SnS_2, soluble.

Arsenic. (84) Heat portion insol. in HCl in closed tube with Na_2Co_3 and KCy. *Black* or *brownish* ring As.

Antimony. (85) HCl sol. treated with Zn gives *black* stain on Pt.

Tin. (86) The metal reduced by Zn is dissolved in HCl. The solution with $HgCl_2$ gives *white* or *gray* precipitate, $\underbrace{Hg_2Cl_2}$.

Gold. (87) Residue dissolved in HCl and HNO_3, evaporated and diluted, gives with $FeSO_4$, $\underbrace{Au}$, *purple, flocculent.*

INSOLUBLE IN AMMONIC SULPHIDE : p. 112, IX.

$\underbrace{HgS}$; $\underbrace{PbS}$; $\underbrace{Bi_2S_3}$; $\underbrace{CuS}$, *black.*—CdS, *yellow.*

Mercury. (89) $\underbrace{HgS}$, insol. in boiling HNO_3.

SOLUBLE IN BOILING NITRIC ACID : p. 112, (88).

PbS ; Bi_2S_3 ; CuS and CdS.

Lead. (90) The solution in HNO_3, after evaporation with H_2SO_4, + water and alcohol, gives $\underbrace{PbSO_4}$, *white, pulverulent.*

Bismuth. (91) The solution in HNO_3, $+ NH_4HO$, gives $\underbrace{Bi(HO)_3}$, *white, flocculent* ; insol. in an excess of NH_4HO.

Copper. (92) The solution in HNO_3, $+ NH_4HO$, gives a *deep blue* solution.

Cadmium. (93) The solution in HNO_3 $+ NH_4HO$ is acidulated with HCl and gives *yellow* CdS with H_2S.

TABLE II.—*Tests for Metals, to be used successively.*

AMMONIC SULPHIDE TEST : p. 115, X.

Metals of Groups III. and IV. in a solution containing NH_4Cl, and neutralized with NH_4HO. **(94)**

INSOLUBLE IN COLD, DILUTE CHLORHYDRIC ACID : p. 116, (96).

$\underbrace{\text{CoS}}$ and $\underbrace{\text{NiS}}$, *black.*

Cobalt. **(97)** The insoluble sulphides give a *blue* borax bead, if much cobalt is present. If the cobalt color is obscured by nickel, it appears after the reduction of the nickel on charcoal.

Nickel. **(98)** When little or no cobalt is present the insoluble sulphides give a *brown* nickel bead. When much cobalt is present, see **(98)**.

SOLUBLE IN COLD, DILUTE CHLORHYDRIC ACID : p. 117 (99).

$\underbrace{\text{FeS}}$, *black ;* $\underbrace{\text{MnS}}$, *flesh-colored ;* $\underbrace{\text{ZnS}}$, white ; $\underbrace{\text{Al}_2(\text{HO})_6}$, *white ;* $\underbrace{\text{Cr}_2(\text{HO})_6}$, *dirty green.*

The HCl solution, freed from H_2S by boiling, and oxidized by HNO_3 (see **99***), is boiled with NaHO in excess (see* **100**).

INSOLUBLE IN SODIC HYDRATE : p. 117 **(100)**.

$\underbrace{(\text{Fe}_2(\text{HO})_6}$; $\underbrace{\text{Mn}(\text{HO})_2}$; $\underbrace{\text{Cr}_2(\text{HO})_6}$.

Iron. **(101)** $Fe_2(HO)_6$ dissolved in HCl, + potassic sulphocyanate, gives a *red* color.

Manganese. **(102)** $Mn(HO)_2$ imparts a *green* color to the mass when fused on platinum foil with Na_2CO_3 and KNO_3.

Chromium. **(102)** $Cr_2(HO)_6$, after oxidation on platinum foil with Na_2CO_3 and KNO_3, boiling with water and filtration, gives a *yellow* (H_2CrO_4) solution on the addition of acetic acid.

SOLUBLE IN SODIC HYDRATE : p. 117 **(100)**.

$Zn(HO)_2$ and $Al_2(HO)_6$.

Zinc. **(103)** The NaHO solution, + H_2S, gives $\underbrace{\text{ZnS}}$, *white, flocculent.*

Aluminium. **(104)** The NaHO solution, + HCl to acid reaction, and then + NH_4HO to alkaline reaction, gives $\underbrace{\text{Al}_2(\text{HO})_6}$, *white, flocculent.*

See p. 119, XI., *for the above tests in the presence of* OXALIC, BORACIC, AND PHOSPHORIC ACIDS.

TESTS FOR METALS OF GROUP II. p. 123, XII.

Add *ammonic hydrate + ammonic carbonate to a solution containing metals of Group II. mixed with ammonic chloride.* $\underbrace{\text{BaCO}_3}$, $\underbrace{\text{SrCO}_3}$, and $\underbrace{\text{CaCO}_3}$, *white.* MAGNESIUM compounds remain in solution.

Barium. **(112)** The precipitate dissolved in dilute chlorhydric acid + $CaSO_4$ gives $\underbrace{\text{BaSO}_4}$, *white, pulverulent.*

Strontium. **(113)** The precipitate dissolved in dilute chlorhydric acid + $CaSO_4$ gives after some time $\underbrace{\text{SrSO}_4}$, *white.*

Calcium. **(115)** The solution **(111)**, free from Ba and Sr, rendered alkaline with $NH_4HO + (NH_4)_2C_2O_4$, gives $\underbrace{\text{CaC}_2\text{O}_4}$, *white, pulverulent.*

Magnesium. **(116)** The solution, freed from Ba, Sr and Ca by treatment with ammonic hydrate and carbonate, + Na_2HPO_4, gives $\underbrace{\text{MgNH}_4\text{PO}_4}$, *white.*

TESTS FOR METALS OF GROUP I. : p. 126, XIII.

Sodium. **(117)** *Yellow* color of flame. (The higher metals must be absent.)

Potassium. **(118)** *Violet* color of flame, seen through a blue glass. (The higher metals must be absent.)

Ammonium. **(119)** Solution, + NaHO, gives NH_3, *ammonia smell.*

TABLE III.—*Tests for Acids. Each test to be applied to a fresh portion of the solution.*

BARIC CHLORIDE TEST: p. 130, XIV.

Acids of Group I., in neutral solution.

$\underbrace{BaCrO_4}$; $\underbrace{BaSO_4}$; $\underbrace{BaSO_3}$; $\underbrace{BaC_2O_4}$; $\underbrace{BaF_2}$; $\underbrace{BaHPO_4}$; $\underbrace{BaB_2O_4}$.

$BaSO_4$ *is the only precipitate insol. in acids.*

Chromic Acid, H_2CrO_4, is reduced to chromic oxide by H_2S. $\underbrace{BaCrO_4}$
(121) gives a *green* borax bead.

Sulphuric Acid, H_2SO_4, + $BaCl_2$ in acid solution gives $\underbrace{BaSO_4}$, *white,*
(123) *pulverulent;* insol. in acids.

Sulphurous Acid, $H_2SO_3 + O$ (oxidation with H_2CrO_4) gives H_2SO_4.
(124) See test for H_2SO_4 above. See also **(46)** and **(75)**.

Oxalic Acid, $H_2C_2O_4$, + $CaSO_4$ in acetic acid solution, gives $\underbrace{CaC_2O_4}$,
(125) *white, pulverulent.*

Fluorhydric Acid, HF, + $CaCl_2$ in acid solution gives $\underbrace{CaF_2}$, *transparent, gelatinous.*
(126)

Fluorine compounds, + H_2SO_4 (concentrated), gives HF, gas which etches glass. See **(44)**.

Phosphoric Acid, H_3PO_4, + ammonic molybdate in nitric acid solution gives phospho-molybdate of ammonium, *yellow crystalline powder.*
(127)

H_2S and Ferro- and Ferricyanhydric acids must be absent. See **(136)** (*a*).

Boracic Acid, H_3BO_3 + when set free by a slight excess of HCl colors
(138) turmeric paper *red*, after moistening and drying.

Arsenious Acid, H_3AsO_3,
Arsenic Acid, H_3AsO_4,
(120)
are precipitated by $BaCl_2$ in neutral solutions; but they should always be detected and separated by HS in the tests for metals. See **(77)**.

ARGENTIC NITRATE TEST: p. 132, XVIII.

Acids of Group II., in acid solution.

$\underbrace{Ag_2S}$, *black;* $\underbrace{AgCl}$, $\underbrace{AgBr}$, $\underbrace{AgCy}$, and $\underbrace{Ag_4(FeCy_6)}$, *white;* $\underbrace{Ag_6(Fe_2Cy_{12})}$, *orange;* AgI, *yellow.*

Sulphydric Acid, H_2S. $\underbrace{Ag_2S}$, *black.* Also
(129) smell of rotten eggs in an acid solution.

Chlorhydric Acid, HCl. $\underbrace{AgCl}$, *white;*
(130) insol. in strong, boiling HNO_3.

Ferro- and ferricyanhydric acids must be absent. See **(136)** (a).

Bromhydric Acid, HBr. $\underbrace{AgBr}$, *white;*
(131) soluble in boiling concentrated HNO_3 and not so soluble in NH_4HO as the chloride.

Iodhydric Acid, HI. $\underbrace{AgI}$, *yellow,* nearly
(132) insoluble in NH_4HO.

Cyanhydric Acid, HCN, + $FeSO_4$, + a few
(133) drops of Fe_2Cl_6, and warmed with NaHO, + HCl, gives the *Prussian blue* color.

Ferro- and ferricyanhydric acids must be absent. See **(136)** (b).

Ferrocyanhydric Acid, H_4FeCy_6, +
(134) Fe_2Cl_6, gives $\underbrace{Fe_4(FeCy_6)_3}$, *Prussian blue.*

Ferricyanhydric Acid, $H_6(Fe_2Cy_{12})$, +
135) $FeSO_4$, gives $\underbrace{Fe_3(Fe_2Cy_{12})}$, *Turnbull's blue.*

ACIDS NOT PRECIPITATED BY ANY METALS.

Acids of Group III.

Chloric Acid, $HClO_3$,
(47) + concentrated, H_2SO_4 gives a *yellow gas.*

Nitric Acid, HNO_3,
(51) + concentrated, H_2SO_4, + $FeSO_4$, gives a *brown* color in an $FeSO_4$ solution.

Acetic Acid, $C_2H_4O_2$,
(52) heated with H_2SO_4 (concentrated), + alcohol, gives the *smell of acetic ether.*

TABLE IV.—EXPLANATION OF SIGNS (*see page* 143); *numbers = degree of solubility*; ∞. = *insoluble*; *s. = soluble*; *s.s. = slightly soluble*; *del. = deliquescent*; *dec. = decomposed*; — = *acids*; + = *alkalies*; *am. = ammonic hydrate*; *am. cl. = ammonic chloride*; *al. = alcohol.*

	Aluminium. Al_2 VI.	Ammonium. NH_4. I.	Antimony. Sb. III. & V.	Barium. Ba. II.	Bismuth Bi. III.	Cadmium. Cd. II.	Calcium Ca. II.	Chromium. Cr_2, VI.	Cobalt. Co. II.	Copper. Cu. II.	Gold. Au. III.	Iron. Fe. II.	Iron. Fe_2. VI.
Acetic Acid, $HC_2H_3O_2$	8,4	del. al. s.	s.	1,4 al. 100.	s.s.	s.	5,6 al. 21.	s.	del.	14,8 al. 13.		s.	s. al. s.
Arsenic Acid, H_3AsO_4	∞. — s.	s.	∞. HCl. s.	s.s. — s.	∞. HCl. s.		∞. — s.	∞. — s.	∞. — s.; am. s.	∞. — s.; am. s.		∞. — s.	∞. — s.
Arsenious Acid, H_3AsO_3. …	∞. — s.	s.	dec. — s.; + s.	s.s. — s.	∞. — s.		4000. — s.		∞. — s.; am. s	∞. — s. + s.		∞. — s.; am. s.	∞. — s
Boracic Acid, H_3BO_3.	s.s. — s.	s.		s.s. — s.	∞. — s.	∞. — s.	s.s. — s.	∞. — s.	s.s. — s.	s.s. — s.		∞. — s.	∞. — s.
Bromhydric Acid, HB.r	del. al. s.	s. al. s.s.	del.	s.	dec.	s.	del.	∞.	del. al. s.	del.	s.	del.	del.
Carbonic Acid, H_2CO_3	∞. — s.	Sesquicarbonate 1,7		12000. — s.	∞. — s.	s. am.cl. s.; — s.	70600. — s.	∞. — s.	∞. — s.	∞. — s.		∞. — s.	∞. — s.
Chlorhydric Acid, HCl.	1,4 al. s.	2,7 al. s.s.	dec. — s.	2,1	dec. — s.	s.	1,4 al. s.	del. al. s.	s. al. s.	1,5 al. s.	1,5	2. al. s.	del. al. s.
Chloric Acid, $HClO_3$.	del. al. s.	s. al. s.		2,7	dec. — s.	del. al. s.	del.		del.	del.		s.	s.
Chromic Acid, H_2CrO_4	∞. — s.	s.		∞. — s.	∞. — s.	s.s.	s.	dec. — s.	∞. — s.	s.s. — s.			∞. — s.
Cyanhydric Acid, HCy	s. al. s.	s. al. s.	∞.	s. al. s.	∞. — s.	s.s. — dec.	s.	∞. — s.	∞. am. s.	∞. — ∞.; + s.	s. al. s.	∞. — s.	
Ferrocyanhydric Acid, $H_4(FeCy_6)$.		s.		s.s. — s.	∞. — ∞.	∞. am. s.; — s.	del. al. ∞.		∞. — ∞.	∞. — ∞.; am. s.		∞. — ∞.	∞. — ∞.
Ferricyanhydric Acid, $H_6(Fe_2Cy_{12})$		s.		s.	∞. — ∞.	∞. am. s.; — s.	del.		∞. — ∞.; am. ∞.	∞. — ∞.; am. s.		∞. — ∞. + dec.	s.
Fluorhydric Acid, HF.	∞.	s. al. s.s.	del.	s.s. — s.	s.	s.s.	27000. — s.s.	s.	s.s. — s.	s.s. — s.		s.s. — s.	s.s. — s.
Iodhydric Acid, HI…	s.	del. al. s.	dec.	del. al. s.	dec.	s. al. s.	del.	s.	del. al. s.	s.	∞.	del.	s.
Nitric Acid, HNO_3…	del. al. 1.	0,5 al. 2.		12,5 al. s.s.	dec. — s.	del. al. s.	2. al. 8.	2. al. s.	del. al. 1.	1. al. s.		s.	s. al. s.
Oxalic Acid, $H_2C_2O_4$.	s.s. — s.	25. al. ∞.	∞. — s.	2600.	∞. — s.	13000. am. s.; — s.	500000. — s.	s.	s.s. — s.; am. s.	∞. — s.		4500. — s.	s.s. — s.
Phosphoric Acid, H_3PO_4. …	∞. — s.	4. al. ∞.	∞.	20500. — s.	∞.	∞. am. cl. s.	∞. — s.	∞. — s.	∞. — s.; am. s.	∞. — s.		∞. — s.	∞. — s.
Sulphuric Acid, H_2SO_4	3. al. ∞.	1,3 al. s.s.	dec. — s.	43000. — ∞.	dec. — s.	s.	380. al. ∞.	s.	44.	3,6		5.	s.
Sulphurous Acid, H_2SO_3 …	∞. — dec.	1. al. s.s.	∞. — dec.	s.s. — dec.	∞. — dec.	s.s. al. ∞.	800.	∞.	s.s. — dec.	s.s. — dec.		s.s. — dec.	s.s. — dec.
Sulphydric Acid, H_2S.	dec.	s. al. s.	∞. — ∞.; + s.	s.	∞. — ∞.	∞.	500.		∞. — dec.	∞. — 8.	∞. — ∞.	∞. — dec.	∞. — dec.
Hydrate, H(HO)…..	∞. — s.; + s.	0,3 al. s.	s.s. — s.; + s.	20.	∞. — s.	∞. am. s.; — s.	584. — s.	∞. — s.; + s.	∞. — s.; am. s.	∞. — s.; am. s.	∞. — s.; + s.	150000. — s.; am. s.s.	∞. — s.

TABLE V.—EXPLANATION OF SIGNS (*see page* 143); *numbers = degree of solubility; ∞. = insoluble; s. = soluble; s.s. = slightly soluble; del. = deliquescent; dec. = decomposed; — = acids; + = alkalies; am. = ammonic hydrate; am. cl. = ammonic chloride; al. = alcohol.*

	Lead. Pb. II.	Magnesium, Mg. II.	Manganese. Mn. II.	Mercu'y Hg$_2$. II.	Mercu'y Hg. II.	Nickel. Ni. II.	Potass'm K. I.	Silver. Ag. I.	Sodium. Na. I.	Strontium. Sr. II.	Tin. Sn. II.	Tin. Sn. IV.	Zinc. Zn. II.
Acetic Acid, $HC_2H_3O_2$	2. al. s.s.	2. al. s.	5. al. s.	769.	2,75	9,6	0,4 del. al. 3.	100.	3,5 al. s.	2,5	s.	s.	4. al. s.
Arsenic Acid, H_3AsO_4	∞. — s.	∞. — s.	∞. — s.	∞. — s.	∞. — s.	∞. — s.; am. s.	del. al. s.	∞. — s.; am. s.	8,4 al. s.	∞. — s.	∞. — s.	∞. — s.	∞. — s.
Arsenious Acid, H_3AsO_3. }	∞. — s.	∞. — s.	∞. — s.	∞. — s.	∞. — s.	∞. — s.; am. s.	del. al. ∞.	∞. — s.; am. s.	s.	s. al. s.s.	∞. — s.	∞. — s.	∞. — s.
Boracic Acid, H_3BO_3.	∞. — s.	s.s. — s.	∞. — s.			∞. — s.	s.	s.s. — s.; am. s.	Borax 12.	s.s.			∞. — s.
Bromhydric Acid, HBr	s.s.	del.	del.	∞.	250.	s.	1,5	∞. am. s.s.	s. al. s.	s.	s.	s.	s. am. s.; — s.
Carbonic Acid, H_2CO_3	50500. — s.	2500. — s.	∞. — s.	∞. — s.	∞. — s.	∞. — s.	0,9 del. al. s.s.	∞. — s.; am. s.	6. al. s.s.	∞. am. cl. s.s.; — s	∞. — s.		∞. — s.
Chlorhydric Acid, HCl.	135. — s.s.	1,8 al. 7.	1,5 al. s.	∞.	18,5 al. 2,5	4.	3. al. s.s.	∞. — ∞.; am. s.	2,6 al. s.s.	del.	s. al. s.	s. al. s.	0,3 al. 1.
Chloric Acid, $HClO_3$.	s. al. s.	del.	s.		4.	del.	16,7 al. 120.	11.	1.	del. al. s.			del.
Chromic Acid, H_2CrO_4	70000. — s.	s.	∞. — s.	∞. — s.	∞. — s.	s.	2. al. ∞.	∞. — s.	s. al. ∞.	s.s. al. ∞.; — s.	∞. — s.	∞. — s.	s.
Cyanhydric Acid, HCy	∞. — s.	s.	∞. — s.	∞.	∞. al. 20.	∞. — ∞.; am. s.	del. al. s.s.	∞. — ∞.; am. s.	s. al. s.s.		∞.		∞. — s.; + s.
Ferrocyanhydric Acid, $H_4(FeCy_6)$. }	∞. — ∞.	s.	∞.	∞. — ∞.	∞.	∞. — ∞.; am. s.	4,5 al. s.s.	∞. — ∞.; am. s	5.	2. al. s.s.			∞. — ∞.; + s.
Ferricyanhydric Acid, $H_6(Fe_2Cy_{12})$ }	s.s. — s.s.	s.	s. — ∞.; am. ∞.			∞. — ∞.; am. s.	2,5 al. s.s.	∞. — ∞.; am. s.	5,5		∞.		∞. — s.; + s.
Fluorhydric Acid, HF.	s.s. — s.	∞. — s.	∞. — s.	∞.	s.s. — s.	s.s. — s.	del. al. s.s.	del.	25.	s.s.	s.	s.	s.s. — s.; am. s.
Iodhydric Acid, HI...	s.s.	del.	del.	∞.	∞.	del.	s.	∞.	s.	s.	s.s.	dec.	del. al. s.
Nitric Acid, HNO_3 ...	1,5	1. al. s.	del.	dec. — s.	dec. — s.	2. al. s.	3,4 al. s.s.	0,44 al. s.	1,14 al. s.s.	5.	dec. — s.	dec.	1. al. s.
Oxalic Acid, $H_2C_2O_4$.	s.s. — s.	2000. — s.; am. cl. s.	3000. — s.	∞. — s.	∞. — s.	∞. — s.; am. s.	3. al. s.s.	s.s. — s.; am. s.	30. al. ∞.	12000.	s.s. — s.	s. — s.; am. s.	∞. — s.; am. s.
Phosphoric Acid, H_3PO_4. } ...	∞. — s.	∞. — s.	∞. — s.	∞. — s.	∞. — s.	∞. — s.	s. al. ∞.	∞. — s.; am. s.	24. al. ∞.	∞. am. cl. s.; — s.	∞. — s.	∞. — s.	∞. — s.; am. s.
Sulphuric Acid, H_2SO_4	13000. — ∞.; al. ∞.	3. al. ∞.	1,5	s.s.	dec. — s.	3.	9,6 al. s.s.	200.	7,4 al. s.s.	6895.	s.	s.	2. al. ∞.
Sulphurous Acid, H_2SO_3 } ...	∞. — dec.	s.s. — dec.	s.s. — dec.		dec. — dec.	∞. — dec.	del. al. s.s.	s.s. — dec.	9.	s.s.			s.s. — dec.
Sulphydric Acid, H_2S.	∞. — ∞.	s.s. — dec.	∞. — dec.	∞. — ∞.	∞. — ∞.	∞. — dec.	del. al. s.	∞. — ∞.	s. al. s.s.	s.	∞. — ∞.; + s.	∞. — ∞.; + s.	∞. — dec.

V1/50